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LECTURE NOTES

FOR

CHEMICAL STUDENTS.

BY

EDWARD FRANKLAND, D.C.L., F.R.S.,

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LECTURE NOTES

FOR

CHEMICAL STUDENTS.

CHAPTER I.

INTRODUCTORY.

DEFINITION.—Chemistry is the science which treats of the composition of all kinds of matter, and of those changes in composition which result from the action, either of different kinds of matter upon each other, or of external forces upon one and the same kind of matter.

SIMPLE AND COMPOUND MATTER.—All kinds of matter are divided into two great classes—*simple substances*, and *compound substances*. A simple substance is one out of which it is impossible to obtain, by any known process, two or more essentially different kinds of matter. A compound substance, on the other hand, is one which can be resolved into two or more simple substances. The simple substances at present known are sixty-two in number, and are termed *elements*. By the combination of these elements with each other, all the infinitely varied forms of terrestrial matter are produced.

MODES OF CHEMICAL ACTION.—Matter undergoes chemical change in five different ways, viz.:—

1st. By the direct combination of elements or compounds with each other.

2nd. By the displacement of one element or group of elements in a body by another element or group of elements.

3rd. By a mutual exchange of elements or groups of elements in two or more bodies.

4th. By the rearrangement of the elements or groups of elements already contained in a body.

5th. By the resolution of a compound into its elements, or into two or more less complex compounds.

Atomic Weight.—Chemists assign to every element a number called its *atomic weight*. This number is not fixed according to any invariable rule, but the following, especially the second, are the chief considerations which serve as guides in its determination:—

1st. The smallest proportion by weight in which the element enters into or is expelled from a chemical compound,—the smallest weight of hydrogen so entering or leaving a chemical compound being taken as unity.

2nd. The weight of the element in the solid condition which, at any given temperature, contains the same amount of heat as seven parts by weight of solid lithium at the same temperature.

3rd. The weight of the element which, in the form of gas or vapour, occupies, under like conditions of temperature and pressure, the same volume as one part by weight of hydrogen.

The atomic weight of a compound is the sum of the atomic weights of its elements.

The atomic weights of the elements are given in the Table at page 6.

Atoms and Molecules.—The proportional amount of any element represented by its atomic weight, as above described, is commonly called an *atom* of that element.

When an element is isolated, or separated from every other kind of matter, its atoms still exist, except in a few cases, in combination with each other. In many instances the atoms of isolated elements are associated in pairs when thus combined. Such an isolated atom or group of atoms constitutes an *elementary molecule*.

The bulk of a molecule, or the *molecular volume* of an element in the gaseous or vaporous condition, is the same as the molecular volume of hydrogen at the same temperature and pressure; and the *molecular weight* of an element is in a large number of cases twice its own atomic weight.

The following is a list of those elements whose molecular volumes have been determined.

Molecules containing of the element

One atom. *Monatomic Molecules.*	Two atoms. *Diatomic Molecules.*	Three atoms. *Triatomic Molecules.*	Four atoms. *Tetratomic Molecules.*	Six atoms. *Hexatomic Molecules.*
Mercury.	Hydrogen.	Oxygen (as Ozone).	Phosphorus.	Sulphur.
Cadmium.	Oxygen.		Arsenic.	
Zinc.	Chlorine.			
	Bromine.			
	Iodine.			
	Fluorine.			
	Nitrogen.			
	Sulphur.			
	Selenium.			

It will be perceived from the above Table that an element may have two distinct molecular weights. This is known to be the case with oxygen and sulphur.

The molecular weight of a compound is, with very few exceptions, identical with its atomic weight. The molecular volume or the space occupied by the combining proportion of a compound is, with very few exceptions, equal to that occupied by two combining proportions, or one molecule, of hydrogen. Hence the law—*equal volumes of all gases and vapours contain, at the same temperature and pressure, an equal number of molecules.*

With very few exceptions, therefore, the molecules of all compounds in the gaseous or vaporous condition, no matter how great may be the aggregate volume of their constituents, occupy, when compared at the same temperature and pressure, one uniform volume, which is exactly the same as that filled by one molecule of hydrogen. Thus

vol.	vol.		vols.
1 of Hydrogen	+ 1 of Chlorine	form	2 of Hydrochloric acid.
1 of Hydrogen	+ 1 of Bromine vapour	„	2 of Hydrobromic acid.
2 of Hydrogen	+ 1 of Sulphur vapour	„	2 of Sulphuretted Hydrogen.
2 of Hydrogen	+ 1 of Oxygen	„	2 of Steam.
3 of Hydrogen	+ 1 of Nitrogen	„	2 of Ammonia.
4 of Hydrogen	+ x of Carbon vapour	„	2 of Marsh-gas.
6 of Hydrogen	+ 1 of Oxygen + $2x$ of Carbon vapour	„	2 of Alcohol vapour.
12 of Hydrogen	+ 1 of Oxygen + $5x$ of Carbon vapour	„	2 of Amylic alcohol vapour.

Chemical Affinity.—The force or power which holds together the elements of a compound is termed *chemical affinity.*

Elements which readily combine with each other, and develop much heat on combination, are said to have a powerful affinity for each other. The elements which thus exhibit towards each other a great affinity are possessed of widely different properties; and when their compounds are decomposed by an electric current, the constituents are evolved at the opposite poles. Those elements which, under such circumstances, make their appearance at the *positive* pole are termed *electro-negative* or *negative* elements, whilst those disengaged at the negative pole are called *electro-positive* or *positive* elements. It must be remembered, however, that the difference between these two classes is one of degree only; they insensibly merge into each other, since the members of both classes exhibit a graduated intensity of the positive or negative quality. Thus potassium is more positive than sodium, and oxygen more negative than sulphur, whilst mercury is negative to sodium but positive to chlorine.

The following eight elements are negative towards the remaining fifty-four elements, which are more or less positive:—

Fluorine.
Chlorine.
Bromine.
Iodine.
Oxygen.
Sulphur.
Selenium.
Tellurium.

Although two positive or two negative elements can combine together chemically, yet their union is rarely attended with such striking phenomena as are manifested when the combination takes place between a positive and a negative element.

CHAPTER II.

CHEMICAL NOMENCLATURE.

The study of every science necessitates an acquaintance with the system of names and peculiar modes of expression which have been found most convenient to denote the materials and to describe the phenomena which form its objects. Such names and modes of expression form the groundwork of the language of every science, upon the right employment of which depend the precision and accuracy of scientific definition.

The nomenclature of a science ought to be distinguished for its clearness and simplicity; but it is by no means easy to secure these conditions in a science like chemistry, where the rapid progress of discovery necessitates the continual addition of new and the frequent alteration of old names. The chemical name of a substance should not only identify and individualize that substance, but it should also express the composition and constitution of the body, if a compound, to which it is applied. The first of these conditions is readily attained; but the second is much more difficult to secure, inasmuch as our ideas of the constitution of chemical compounds—of the mode in which they are built up as it were—require frequent modification. On this account all attempts to frame a perfectly consistent system of chemical nomenclature have hitherto been only partially successful.

It has been already mentioned that the number of elements at present known is sixty-two. These have received the names given in the following Table, in which the twenty-two most important elements are distinguished by the largest type, those next in importance by medium type, whilst the names of elements which are either of rare occurrence, or of which our knowledge is yet very imperfect, are printed in the smallest type.

Name.	Symbol.	Atomic weight.	Name.	Symbol.	Atomic weight.
ALUMINIUM	Al	27·5	Molybdenum	Mo	92
ANTIMONY	Sb	122	NICKEL	Ni	58·8
ARSENIC	As	75	Niobium	Nb	97·6
BARIUM	Ba	137	**NITROGEN**	N	14
BISMUTH	Bi	208	Osmium	Os	199
BORON	B	11	**OXYGEN**	O	16
BROMINE	Br	80	PALLADIUM	Pd	106·5
Cadmium	Cd	112	**PHOSPHORUS**	P	31
Cæsium	Cs	133	PLATINUM	Pt	197·4
CALCIUM	Ca	40	**POTASSIUM**	K	39
CARBON	C	12	RHODIUM	Rh	104
Cerium	Ce	92	Rubidium	Rb	85·5
CHLORINE	Cl	35·5	Ruthenium	Ru	104
CHROMIUM	Cr	52·5	Selenium	Se	79
COBALT	Co	58·8	**SILICON**	Si	28·5
COPPER	Cu	63·5	**SILVER**	Ag	108
Didymium	D	96	**SODIUM**	Na	23
FLUORINE	F	19	STRONTIUM	Sr	87·5
Glucinum	G	14	**SULPHUR**	S	32
GOLD	Au	196·7	Tantalum	Ta	137·5
HYDROGEN	H	1	Tellurium	Te	128
Indium	In	113·4	Thallium	Tl	204
IODINE	I	127	Thorium	Th	231·5
IRIDIUM	Ir	198	TIN	Sn	118
IRON	Fe	56	TITANIUM	Ti	50
Lanthanum	L	92	TUNGSTEN	W	184
LEAD	Pb	207	URANIUM	U	120
Lithium	Li	7	Vanadium	V	51·2
MAGNESIUM	Mg	24	Yttrium	Y	68
MANGANESE	Mn	55	**ZINC**	Zn	65
MERCURY	Hg	200	Zirconium	Zr	90

These elementary substances have been long divided into two great classes—*metals* and *non-metals*, the latter being also sometimes termed *metalloids*. The metals are by far the more numerous, the non-metals comprising only the following thirteen elements—Boron, Bromine, Carbon, Chlorine, Fluorine, Hydrogen, Iodine, Nitrogen, Oxygen, Phosphorus, Selenium, Silicon, Sulphur.

The names of the elements can scarcely be said to have been given according to any rule; many of them are derived from some prominent property of the bodies themselves, whilst others have a mythological origin. An attempt has been made to distinguish the metals by the termination *um*, as potassium, sodium, &c.; but the common metals, such as gold, copper, and iron, still retain their original names; and one substance, selenium, which at the time of its discovery was regarded as a metal, has had no change made in its name, although further research has divested it of all metallic attributes. An important group of electro-negative non-metals—fluorine, chlorine, bromine, and iodine—have received the termination *ine*; three are distinguished by the terminal syllable *on*, viz. carbon, silicon, and boron; and three others have *gen* for their final syllable, viz. oxygen, hydrogen, and nitrogen, these last names being derived from Greek words denoting the property possessed by these elements of generating respectively acid, water, and nitre.

When two elementary bodies unite together, they form a chemical compound of the first order, to which the name *binary compound* has been applied. The names of these compounds are formed from those of their constituents, the name of the positive constituent or some abbreviation thereof, with the terminal *ic*, preceding that of the negative constituent, which is made to terminate in *ide*, thus:—

Potassium	and	Sulphur	form	Potassic sulphide.
Sodium	„	Oxygen	„	Sodic oxide.
Silver	„	Chlorine	„	Argentic chloride.
Zinc	„	Iodine	„	Zincic iodide.
Calcium	„	Chlorine	„	Calcic chloride.

But the same elements frequently form with each other two compounds, in which case the one which contains the smaller proportion of the negative element is distinguished by changing the terminal syllable of the name of its positive constituent into *ous*, the terminal *ic* being retained for the compound containing the larger proportion of the negative element. Thus,

One atom of tin and two atoms of chlorine form Stannous chloride.
One atom of tin and four atoms of chlorine form Stannic chloride.

Sometimes, however, the same elements form with each other more than two compounds. In these cases the prefixes *hypo* and *per* are employed as marks of distinction; but their use is very rarely required.

If a binary compound contains oxygen, and forms an acid when made to unite with water, or a salt when added to a base, it is termed an *anhydride* or *anhydrous acid.* Thus,

One atom of carbon and two atoms of oxygen form carbonic anhydride.
Two atoms of nitrogen and five atoms of oxygen form nitric anhydride.
Two atoms of nitrogen and three atoms of oxygen form nitrous anhydride.
One atom of sulphur and three atoms of oxygen form sulphuric anhydride.
One atom of sulphur and two atoms of oxygen form sulphurous anhydride.

In the following cases, the systematic names have not displaced the trivial and irregular names used for the same substances:—

Systematic Name.	Trivial or irregular name.
Hydric oxide	Water.
Hydric sulphide	Sulphuretted hydrogen.
Hydric selenide	Seleniuretted hydrogen.
Hydric telluride	Telluretted hydrogen.
Hydric chloride	Hydrochloric acid.
Hydric bromide	Hydrobromic acid.
Hydric iodide	Hydriodic acid.
Hydric fluoride	Hydrofluoric acid.
Hydric carbide	Marsh-gas or light carburetted hydrogen.
Hydric nitride	Ammonia.
Hydric phosphide	Phosphuretted hydrogen.
Hydric arsenide	Arsenuretted hydrogen
Hydric antimonide	Antimonuretted hydrogen.

The term *acid* was originally applied only to substances possessing a sour taste like vinegar; but analogy has necessitated the application of the same name to a large number of compounds which have not this property. In the modern acceptation of the name, an acid may be defined as a compound containing one or more atoms of hydrogen, which become displaced by a metal when the latter is presented to the compound in the form of a hydrate. The hydrogen capable of being so displaced may be conveniently termed *displaceable hydrogen.* An acid containing one such atom of hydrogen is said to be *monobasic*, two such atoms *dibasic*, &c. Acids of a greater basicity than unity are frequently termed *polybasic acids.*

Thus nitric acid gives, with sodic hydrate, sodic nitrate:

$$\underset{\text{Nitric acid.}}{NO_3H} + \underset{\text{Sodic hydrate.}}{ONaH} = \underset{\text{Sodic nitrate.}}{NO_3Na} + \underset{\text{Water.}}{OH_2}.$$

Sulphuric acid gives, with potassic hydrate, potassic sulphate:

$$\underset{\text{Sulphuric acid.}}{SO_4H_2} + \underset{\text{Potassic hydrate.}}{2OKH} = \underset{\text{Potassic sulphate.}}{SO_4K_2} + \underset{\text{Water.}}{2OH_2}.$$

And hydrochloric acid gives, with potassic hydrate, potassic chloride:

$$\underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{HCl} + \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{OKH} = \underset{\substack{\text{Potassic}\\\text{chloride.}}}{KCl} + \underset{\text{Water.}}{OH_2}.$$

When an acid contains oxygen, its name is generally formed by adding the terminal *ic* either to the name of the element with which the oxygen is united, or to an abbreviation of that name; thus sulphur forms, with oxygen, sulphuric acid; nitrogen, nitric acid; and phosphorus, phosphoric acid. But it frequently happens that the same element forms two acids with oxygen; and when this occurs, the acid containing the larger amount of oxygen receives the terminal syllable *ic*, whilst that containing less oxygen is made to end in *ous.* Thus we have sulphurous acid, nitrous acid, and phosphorous acid, each containing a smaller proportion of oxygen than that necessary to form respectively sulphuric, nitric, and phosphoric acids.

In some instances, however, the same element forms more than two acids with oxygen, in which case one of the two Greek words *hypo*, under, and *hyper*, over, is prefixed to the name of the acid. Thus an acid of sulphur containing less oxygen than sulphurous acid is termed hyposulphurous acid; and another acid of the same element containing, in proportion to sulphur, more oxygen than sulphurous acid and less than sulphuric, might be named either hypersulphurous acid, or hyposulphuric acid; but the latter term has been universally adopted. The prefix *per* is frequently substituted for *hyper*; thus in the case of chlorine, which forms the following four acids with oxygen, viz. hypochlorous acid, chlorous acid, chloric acid, and hyperchloric acid, the latter is generally named perchloric acid; but *per* can only be used as a prefix to the acid containing the *largest proportion of oxygen*.

Some acids do not contain oxygen amongst their constituents, but consist of sulphur or hydrogen united with other elements. This peculiarity of composition is expressed in their nomenclature by the prefixes *sulpho* or *sulph*, and *hydro* or *hydr*: thus sulpharsenic acid and sulphostannic acid denote acids composed respectively of sulphur, hydrogen, and arsenic; and sulphur, hydrogen, and tin; whilst the names hydrochloric acid and hydriodic acid are given to acids composed, the first of hydrogen and chlorine, and the second of hydrogen and iodine. The terminals *ous* and *ic* are also applied to these acids in exactly the same manner as to the oxygen acids: thus we have sulpharsenious and sulpharsenic acid, the latter containing a larger proportion of sulphur than the former; but the application of the first of these terminals has not hitherto been found necessary in the case of hydrogen acids, since no element has yet been observed to form more than one acid with hydrogen.

The term *anhydride* or *anhydrous acid* is applied to the residue obtained by the abstraction (in combination with oxygen as water) of all the displaceable hydrogen from one or two molecules of an oxygen acid. Thus,

$$\underset{\text{Sulphuric acid.}}{SO_4H_2} - \underset{\text{Water.}}{OH_2} = \underset{\text{Sulphuric anhydride.}}{SO_3};$$

$$\underset{\text{Nitric acid.}}{2NO_3H} - \underset{\text{Water.}}{OH_2} = \underset{\text{Nitric anhydride.}}{N_2O_5}.$$

The term ***base*** is applied to three classes of compounds, all of which are converted into salts by the action of acids. These are—

1st. Certain compounds of metals with oxygen, such as sodic oxide (Na_2O), zincic oxide (ZnO), &c.

2nd. Certain compounds of metals with the compound radical hydroxyl (HO), such as sodic hydrate $\left(Na(HO)\right)$, zincic hydrate $\left(Zn(HO)_2\right)$, &c.

3rd. Certain compounds of nitrogen, phosphorus, arsenic, and antimony, such as ammonia (NH_3).

There are also a few organic compounds to which the name base is sometimes given, but which are not included in the above classes; it is, however, unnecessary further to allude to them here.

The bases of the first class are named in accordance with the rules already given for compounds of two elements. The following bases, however, still retain their irregular names:—

Systematic names.	Irregular names.
Baric oxide	Baryta.
Strontic oxide	Strontia.
Calcic oxide	Lime.
Magnesic oxide	Magnesia.
Aluminic oxide	Alumina.
Glucinic oxide	Glucina.
Zirconic oxide	Zirconia.

The names of the bases belonging to the second class are formed by changing the terminal syllable of the name of the metal into *ic* or *ous*, and the word hydroxyl into *hydrate.* Thus cæsium and hydroxyl form cæsic hydrate $\left(Cs(HO)\right)$; barium and hydroxyl, baric hydrate $\left(Ba(HO)_2\right)$; and iron and hydroxyl, ferric hydrate $\left(Fe_2(HO)_6\right)$.

A few of these bases have trivial or irregular names, which are almost invariably used instead of the systematic names:—

Systematic names.	Irregular names.
Potassic hydrate	Potash.
Sodic hydrate	Soda.
Lithic hydrate	Lithia.

The bases of the third class are distinguished by the terminal syllable *ine*, except nitrine, (NH_3), which retains its trivial name ammonia. These bases belong almost exclusively to the department of organic chemistry, and their nomenclature could not be advantageously discussed here.

It has been already mentioned that by the mutual action of an acid and a base upon each other, a *salt* is produced. If the salt be free from oxygen and sulphur, like common salt, ($NaCl$), it is termed a *haloid salt* ; if it contain oxygen, it is termed an *oxysalt* ; and if this oxygen be replaced by sulphur, it is distinguished as a *sulphosalt*.

The haloid salts are named according to the rules for binary compounds above given, thus :

Name.	Formula.
Sodic chloride	$NaCl$.
Calcic iodide	CaI_2.
Ferrous bromide	$FeBr_2$.
Ferric bromide	Fe_2Br_6.

Oxysalts are divided into *normal*, *acid*, and *basic*.

A normal salt is one in which the displaceable hydrogen of the acid (see page 9) *is all exchanged for an equivalent amount of a metal or of a positive compound radical.*

The following examples will serve to illustrate this definition of a normal, or, as it is sometimes incorrectly called, a *neutral* salt, the displaceable atoms of hydrogen in the acid, and the metal by which they have been displaced in the salt, being printed in italics :—

Acid.		Normal salt.	
Nitric acid	$NO_3\mathit{H}$	Sodic nitrate	$NO_3\mathit{Na}$.
		Calcic nitrate	$(NO_3)_2\mathit{Ca}''$.
Sulphuric acid	$SO_4\mathit{H}_2$	Potassic sulphate	
		Calcic sulphate	$SO_4\mathit{Ca}''$.
Phosphoric acid	$PO_4\mathit{H}_3$	Potassic phosphate.	
		Calcic phosphate.........	$(PO_4)_2\mathit{Ca}''_3$.
Hypophosphorous acid..	$PO_2H_2\mathit{H}$ ...	Sodic hypophosphite ...	$PO_2H_2\mathit{Na}$.
Phosphorous acid	$PO_3H\mathit{H}_2$ ...	Potassic phosphite	$PO_3H\mathit{K}_2$.

Acid.		Normal salt.	
Metaphosphoric acid ...	PO_3H	Lithic metaphosphate...	PO_3Li.
Pyrophosphoric acid ...	$P_2O_7H_4$......	Calcic pyrophosphate...	$P_2O_7Ca''_2$.
Nordhausen sulphuric acid	$S_2O_7H_2$......	Sodic bisulphate	$S_2O_7Na_2$.
Unknown acid	$Cr_2O_7H_2$ ...	Potassic bichromate ...	$Cr_2O_7K_2$.

An acid salt is one in which the displaceable hydrogen of the acid is only partially exchanged for a metal or positive compound radical.

The following examples illustrate the constitution and nomenclature of these salts :—

Acid.		Acid salt.	
Sulphuric acid ...	SO_4H_2	Hydric sodic sulphate	SO_4HNa.
Carbonic acid ...	CO_3H_2?	Hydric potassic carbonate...	CO_3HK.
Phosphoric acid...	PO_4H_3	Hydric disodic phosphate...	PO_4HNa_2.
		Dihydric sodic phosphate ...	PO_4H_2Na.
		Microcosmic salt .... (Hydric ammonic sodic	$PO_4H(NH_4)Na$.

Acid salts are produced almost exclusively from polybasic acids.

When the number of bonds* of the metal or compound positive radical contained in a salt exceeds the number of atoms of displaceable hydrogen in the acid, the compound is usually termed a basic salt—as, for instance,

Acid.		Basic salt.	
Carbonic acid	CO_3H_2	Malachite	$CO_5H_2Cu''_2$.
		Blue cupric carbonate	$C_2O_8H_2Cu''_3$.
Sulphuric acid.........	SO_4H_2	Tribasic cupric sulphate	$SO_8H_4Cu''_3$.
		Turpeth mineral	$SO_6Hg''_3$.

These and most, if not all, other basic salts do not differ essentially in their constitution from the normal and acid salts. This will be seen from the arrangement of their atoms given under the different metals entering into their composition.

The nomenclature of organic bodies is founded upon the same principles as that of inorganic compounds; but its discussion could not be conveniently introduced here.

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* For an explanation of this term see Chap. III. p. 18.

CHAPTER III.

CHEMICAL NOTATION.

SYMBOLIC NOTATION.—Every element is represented by a symbol, which is frequently the initial letter of the name of the element; but as in some cases the names of two or more elements begin with the same letter, it is necessary to distinguish them by the use of a second letter in small type, which is either the second letter of the word, or some other letter prominently heard in its pronunciation; thus carbon, cadmium, cobalt, and cerium all begin with the same letter; but they are distinguished by the symbols C, Cd, Co, and Ce. In the use of the single letters, the non-metallic elements have the preference; thus oxygen, hydrogen, nitrogen, sulphur, phosphorus, boron, carbon, iodine, and fluorine are expressed by the single letters O, H, N, S, P, B, C, I, and F; whilst the metals osmium, mercury, nickel, strontium, platinum, bismuth, cobalt, iridium, and iron are symbolized by two letters each; thus Os, Hg (hydrargyrum), Ni, Sr, Pt, Bi. Co, Ir, and Fe (ferrum). In the selection of the single letter for other cases, preference is given to the most important element; thus sulphur, selenium, and silicon are all non-metallic elements, beginning with the same letter; but sulphur being the most important, the single letter S is assigned to it, whilst selenium and silicon are denoted respectively by Se and Si.

The symbols of compounds are formed by the simple juxtaposition of the symbols of their constituent elements. Such a group of two or more symbols is termed a *chemical formula*. Thus:

Argentic chloride	·AgCl.
Zincic oxide	ZnO.

The symbols not only represent the elements for which they are used, but they also denote a certain definite proportion by weight of each element; the formula HCl, for instance, does not merely denote a compound of hydrogen and chlorine, but it signifies a molecule of that compound containing one atom

(1 part by weight) of hydrogen, and one atom (35·5 parts by weight) of chlorine. When, therefore, the molecule of a compound contains more than one atom or combining proportion of any element, it is necessary to express such fact in the formula: this is done by the use of a coefficient placed after the symbol of the element:

Zincic chloride	$ZnCl_2$.
Ferric chloride	Fe_2Cl_6.
Stannous chloride	$SnCl_2$.
Stannic chloride	$SnCl_4$.

When it is necessary to denote two or more molecules of any compound, a large figure is placed before the formula of the compound; such a figure then affects every symbol in that formula: thus $3SO_4H_2$ means three molecules of the compound SO_4H_2.

The changes which occur during chemical action are expressed by equations, in which the symbols of the elements or compounds, as they exist before the change, are placed on the left, and those which result from the reaction on the right. Thus, taking an example from each of the five kinds of chemical action before mentioned, we have

$$(1)\qquad \underset{\text{Zinc.}}{Zn} + \underset{\text{Chlorine.}}{Cl_2} = \underset{\text{Zincic chloride.}}{ZnCl_2}.$$

$$(2)\qquad \underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Zinc.}}{Zn} = \underset{\text{Zincic chloride.}}{ZnCl_2} + \underset{\text{Hydrogen.}}{H_2}.$$

$$(3)\qquad \underset{\text{Cupric sulphate.}}{SO_4Cu} + \underset{\text{Baric nitrate.}}{(NO_3)_2Ba} = \underset{\text{Baric sulphate.}}{SO_4Ba} + \underset{\text{Cupric nitrate.}}{(NO_3)_2Cu}.$$

$$(4)\qquad \underset{\text{Ammonic cyanate.}}{(CN)O(NH_4)} = \underset{\text{Urea.}}{N_2H_4(CO)}.$$

$$(5)\qquad \underset{\text{Water.}}{2OH_2} = \underset{\text{Oxygen.}}{O_2} + \underset{\text{Hydrogen.}}{2H_2}.$$

The sign $+$, as seen from the foregoing examples, is placed between the formulæ of the molecules of the different substances which are brought into contact before the reaction, and of those which result from the change. This sign must never be used to connect together the constituents of one and the same chemical compound.

The sign — is only very rarely used in chemical notation, but when employed it has the ordinary signification of abstraction; thus,

$$\underset{\text{Sulphuric acid.}}{SO_4H_2} - \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphuric anhydride.}}{SO_3}.$$

Use of the bracket.—The bracket has been employed in various senses in chemical formulæ; but in the following pages it is used in notation for one purpose only, viz. for expressing chemical combination between two or more elements which are placed perpendicularly with regard to each other and next to the bracket in a formula. Thus in the following cases,

$$\begin{array}{ccc} \text{I.} & \text{II.} & \text{III.} \\ CH_3 & \left\{\begin{array}{l} CH_3 \\ O \\ CH_3 \end{array}\right. & \left.\begin{array}{r} NO_2O \\ Ba \\ NO_2O \end{array}\right\} \end{array}$$

the formula No. I. signifies that two atoms of carbon are directly united with each other, No. II. that two atoms of carbon are linked together, as it were, by an atom of oxygen, the latter being united to both carbon atoms; whilst in like manner No. III. indicates that one atom of oxygen in the formula of the upper line is linked to another atom of oxygen in the formula of the lower line, by an atom of barium.

Use of thick letters.—As a rule, the formulæ in this book are so written as to denote that the element represented by the first symbol of a formula is directly united with all the active bonds of the other elements or compound radicals following upon the same line: thus the formula $\mathbf{S}O_2Ho_2$ (sulphuric acid) signifies that the hexad atom of sulphur is combined with the four bonds of the two atoms of oxygen, and also with the two bonds of the two semi-molecules of hydroxyl. Such a formula is termed a *constitutional formula**.

Occasionally, however, owing to the atomic arrangement of a compound not being known, its formula cannot be written according to this rule; and in order to prevent such *molecular* or *empirical formulæ*[*] from being mistaken for constitutional formulæ, the first symbol of a constitutional formula will always be printed in thick type. It deserves also to be mentioned that, as a rule, the element having the greatest number of bonds will occupy this prominent position. Thus,

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* For further information on this subject see ATOMICITY OF ELEMENTS and COMPOUND RADICALS, pages 17 and 26.

Sulphuric acid	$\mathbf{S}O_2Ho_2$.
Water	$\mathbf{O}H_2$.
Nitric acid	$\mathbf{N}O_2Ho$.
Microcosmic salt	$\mathbf{P}OHoAmoNao$.

ATOMICITY OF ELEMENTS.—It has been already stated that the atomic weight of an element is the smallest proportion by weight in which that element enters into or is expelled from a chemical compound. The atoms of the various elements, the relative weights of which are thus expressed, exhibit very different values in chemical reactions. Thus an atom of zinc is equivalent to two atoms of hydrogen; for, when zinc is brought into contact with steam at a high temperature, one atom of zinc expels from the steam two atoms of hydrogen and occupies their place—thus,

$$\underset{\text{Water.}}{\mathbf{O}H_2} + Zn = \underset{\text{Zincic oxide.}}{\mathbf{O}Zn} + H_2.$$

Again, when zincic oxide is brought into contact with hydro-

---

[*] A *molecular formula*, sometimes called *rational*, is one in which the atomic composition of a molecule is expressed, but without reference to the manner in which the elements are combined amongst themselves. An *empirical formula* merely expresses, by the smallest integers, the proportional number of atoms of each element entering into the composition of a compound. Thus the three formulæ of ferric hydrate are written:—

Empirical formula	FeH_3O_3·
Molecular "	$Fe_2H_6O_6$.
Constitutional "	$\mathbf{F}e_2Ho_6$.

Thus both constitutional and rational formulæ are essentially molecular formulæ, whilst empirical formulæ afford no indication of the number of atoms entering into the composition of a molecule; they are, in fact, only used to represent bodies the molecular weights of which are unknown.

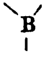
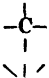
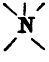
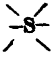
chloric acid, the place of the zinc becomes once more occupied by hydrogen, but two atoms of hydrogen are found to be necessary to take the place of one atom of zinc :

$$\underset{\text{Zincic oxide.}}{\mathbf{O}Zn} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Zincic chloride.}}{\mathbf{Zn}Cl_2} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

In like manner one atom of boron can be substituted for three atoms of hydrogen, one atom of carbon for four, one of nitrogen for five, and one atom of sulphur for no less than six atoms of hydrogen.

This combining value of the elementary atoms is termed their *atomicity*, *equivalence*, or *atom-fixing power* ; and an element with an atom-fixing power equal to that of one atom of hydrogen is termed a *monad*, with twice that power a *dyad*, with thrice a *triad*, with quadruple a *tetrad*, with quintuple a *pentad*, and, with an atom-fixing power equal to six times that of hydrogen, a *hexad*.

To avoid any speculation as to the nature of the tie which enables an element thus to attach to itself one or more atoms of other elements, I have named each unit of atom-fixing power a *bond*,—a term which involves no hypothesis as to the nature of the connexion. A monad element has obviously only one such bond ; a dyad, like zinc, two; a triad, like boron, three, and so on. The number of bonds possessed by an elementary atom may be graphically represented by lines in the following manner :—

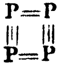
Hydrogen	H–
Zinc	–Zn–
Boron..............................	
Carbon	
Nitrogen...........................	
Sulphur	

In symbolic notation, the same idea is conveyed by the use of dashes and Roman numerals placed above and to the right of the symbol of the element; thus,

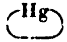
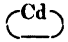
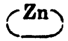
Hydrogen.........	H',	Carbon............	C^{iv},
Zinc	Zn'',	Nitrogen	N^{v},
Boron	B''',	Sulphur	S^{vi}.

Elements with an odd number of bonds are termed ***perissads***, whilst those with an even number are named ***artiads***.

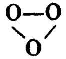
No element, either alone or in combination, can exist with any of its bonds disconnected; hence ***the molecules of all elements with an odd number of bonds are generally diatomic, and always polyatomic***; that is, they contain two or more atoms of the element united together. Thus,

	Symbolic.	Graphic.
Hydrogen	H_2	$H—H$
Chlorine	Cl_2	$Cl—Cl$
Nitrogen	N^{v}_2	$N\equiv N$
Phosphorus.........	P^{v}_4	

An element with an even number of bonds can exist as a monatomic molecule, its own bonds satisfying each other. Thus,

	Symbolic.	Graphic.
Mercury	Hg''	
Cadmium	Cd''	
Zinc	Zn''	

It is nevertheless obvious that such an element may also exist as a polyatomic molecule. Oxygen furnishes us with an example of this; for in its ordinary condition it is a diatomic molecule, and in the allotropic form of ozone, a triatomic molecule:

	Symbolic.	Graphic.
Oxygen	O''_2	O=O
Ozone	O''_3	

In order to avoid the unnecessary use of atomicity-marks in symbolic notation, I shall never attach them to a monad, or to oxygen, which, it must be remembered, is always a dyad. Neither will the atomicity-coefficient be attached to the tetrad element carbon, in the formulæ of organic bodies, unless this element plays the part of a dyad—an occurrence of extreme rarity. When not otherwise marked, therefore, carbon must always be understood to be a tetrad.

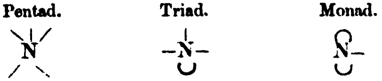
It will also, as a rule, be unnecessary to mark the atomicity of the elements which are expressed by symbols in thick type, because their atomicity is clearly indicated by the sum of the atomicities of the elements or compound radicals placed to their right, or connected with them perpendicularly by a bracket. Thus in the formula

$$\left\{\begin{matrix}\mathbf{C}Cl_3\\ \mathbf{C}Cl_3\end{matrix}\right.,$$

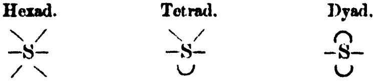
each atom of carbon is united with three atoms of the monad chlorine, whilst the bracket indicates that the two atoms of carbon are also united by one bond of each, thus denoting **C** to be a tetrad element.

From what has just been said with regard to carbon, it is evident that the atomicity of an element is, apparently at least, not a fixed and invariable quantity: thus nitrogen is sometimes equivalent to five atoms of hydrogen, as in ammonic chloride, ($\mathbf{N}^{v}H_4Cl$), sometimes to three atoms, as in ammonia ($\mathbf{N}'''H_3$), and sometimes to only one atom, as in nitrous oxide ($\mathbf{O}N_2$). But it is found that this variation in atomicity always takes place by the disappearance or development of an even number of bonds: thus nitrogen is either a pentad, a triad, or a monad; phosphorus and arsenic, either pentads or triads; carbon and tin, either tetrads or dyads; and sulphur, selenium, and tellurium, either hexads, tetrads, or dyads.

These remarkable facts can be explained by a very simple and obvious assumption, viz. that *one or more pairs of bonds belonging to the atom of an element can unite and, having saturated each other, become, as it were, latent.* Thus the pentad element nitrogen becomes a triad when one pair of its bonds becomes latent, and a monad when two pairs, by combination with each other, are, in like manner, rendered latent,—conditions which may be graphically represented thus:—



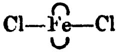
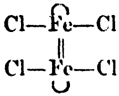
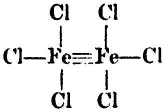
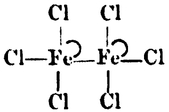
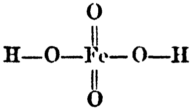
And in the case of sulphur:—



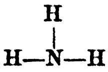
Adopting this hypothesis, it will be convenient to distinguish the maximum number of bonds of an element as its *absolute atomicity*, the number of bonds united together as its *latent atomicity*, and the number of bonds actually engaged in linking it with the other elements of a compound as its *active atomicity*. The sum of the active and latent atomicity of any element must evidently always be equal to the absolute atomicity. Thus in sulphuric acid ($\mathbf{S}^{\text{vi}}O_2Ho_2$) the absolute and active atomicities are both =VI, therefore the latent atomicity =0. In sulphurous acid ($''S^{\text{iv}}OHo_2$) the active atomicity =IV, and consequently the latent = VI − IV = II; whilst in sulphuretted hydrogen ($^{\text{iv}}\mathbf{S}''H_2$) the active and latent atomicities are respectively II and IV.

The apparent exceptions to this hypothesis nearly all disappear on investigation: thus iron, which is a dyad in ferrous compounds (as $\mathbf{Fe}Cl_2$), a tetrad in cubical pyrites ($\mathbf{Fe}S''_2$), and a hexad in ferric acid ($\mathbf{Fe}O_2Ho_2$), is apparently a triad in ferric chloride ($\mathbf{Fe}Cl_3$); but the vapour-density of ferric chloride shows that its formula must be doubled—that, in fact, the two

atoms of the hypothetical molecule of iron (Fe_2) have not been completely separated. The formulæ of the ferrous and ferric chlorides and of ferric acid then become

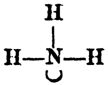
	Symbolic.	Graphic.
Ferrous chloride...	${}^{iv}\mathbf{Fe}''Cl_2$.	 $\overset{\frown}{\underset{\smile}{Cl—Fe—Cl}}$
or	$''\begin{cases} ''\mathbf{Fe}Cl_2 \\ ''\mathbf{Fe}Cl_2 \end{cases}$	 $\begin{array}{c} Cl—\overset{\frown}{Fe}—Cl \\ \Vert \\ Cl—\underset{\smile}{Fe}—Cl \end{array}$
Ferric chloride ...	$''\mathbf{Fe}'''_2Cl_6$.	 $\begin{array}{c} Cl \quad Cl \\ \vert \quad\ \vert \\ Cl—Fe\equiv Fe—Cl \\ \vert \quad\ \vert \\ Cl \quad Cl \end{array}$
or	$\begin{cases} ''\mathbf{Fe}Cl_3 \\ ''\mathbf{Fe}Cl_3 \end{cases}$	 $\begin{array}{c} Cl \quad Cl \\ \vert \quad\ \vert \\ Cl—Fe—Fe—Cl \\ \vert \quad\ \vert \\ Cl \quad Cl \end{array}$
Ferric acid	$\mathbf{Fe}^{vi}O_2Ho_2$.	 $\begin{array}{c} O \\ \Vert \\ H—O—Fe—O—H \\ \Vert \\ O \end{array}$

It will be remarked that the number of bonds supposed to be combined with each other in the atom of iron in ferrous chloride is expressed in one of the above formulæ by the atomicity numeral IV placed to the left of the symbol, whilst the analogous union of three bonds of each atom of iron in ferric chloride is expressed by the three dashes ''' to the left of the symbol $\mathbf{Fe}_2$. I shall not, however, use these coefficients of latent atomicity in the case of the single atom of an element, the student being supposed to have made himself acquainted with the absolute atomicity of every element as expressed in the Table at page 32. For a similar reason, it will also rarely be necessary to express the same idea in graphic notation: thus, for instance, ammonia will be drawn



$$\begin{array}{c} H \\ | \\ H\text{—}N\text{—}H \end{array}$$

and not



$$\begin{array}{c} H \\ | \\ H\text{—}N\text{—}H \\ \smile \end{array}$$

It will be necessary, however, to employ these coefficients in symbolic formulæ where two or more atoms of the same element are joined together under such circumstances that the number of bonds uniting them cannot be found by subtracting the coefficient of active atomicity from the absolute atomicity of the element; as in hydric persulphide ($'S'_2H_2$), for instance, which might otherwise be viewed as $'''S'_2H_2$, or $'S'_2H_2$.

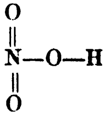
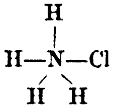
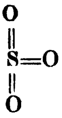
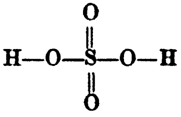
In rare cases, in which oxygen links together two elements or radicals in the same line of a formula, a hyphen is placed before and after the symbol O, thus:—

$$\left\{\begin{array}{l} CH_2\text{-}O\text{-}\mathbf{C}MeO \\ CH_2\text{-}O\text{-}\mathbf{C}MeO\text{·} \end{array}\right.$$

Diacetic glycol.

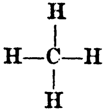
GRAPHIC NOTATION.—This mode of notation, although far too cumbrous for general use, is invaluable for clearly showing the arrangement of the individual atoms of a chemical compound. It is true that it expresses nothing more than the symbolic notation of the same compound, if the latter be written and understood as above described; nevertheless the graphic form affords most important assistance, both in fixing upon the mind the true meaning of symbolic formulæ, and also in making comparatively easy of comprehension the probable internal arrangement of the very complex molecules frequently met with both in mineral and organic compounds. It is also of especial value in rendering strikingly evident the causes of isomerism in organic bodies; and it is now almost universally employed by chemists in describing the results of their new discoveries.

Graphic notation, like the above method of symbolic notation, is founded essentially upon the doctrine of atomicity, and consists in representing, graphically, the mode in which every bond in a chemical compound is disposed of. Inasmuch, however, as the principles involved are precisely the same as those already described under the heads of SYMBOLIC NOTATION and ATOMICITY OF ELEMENTS, it is unnecessary here to do more than give the following comparative examples of symbolic and graphic formulæ:—

	Symbolic.	Graphic.
Water	$\mathbf{O}\mathrm{H}_2$.	H—O—H
Nitric acid	$\mathbf{N}\mathrm{O}_2\mathrm{Ho}$.	 $\begin{array}{l} \mathrm{O} \\ \| \\ \mathrm{N{-}O{-}H} \\ \| \\ \mathrm{O} \end{array}$
Ammonic chloride	$\mathbf{N}\mathrm{H}_4\mathrm{Cl}$.	
Sulphuric anhydride	$\mathbf{S}\mathrm{O}_3$.	 $\begin{array}{l} \mathrm{O} \\ \| \\ \mathrm{S{=}O} \\ \| \\ \mathrm{O} \end{array}$
Sulphuric acid	$\mathbf{S}\mathrm{O}_2\mathrm{Ho}_2$.	 $\begin{array}{c} \mathrm{O} \\ \| \\ \mathrm{H{-}O{-}S{-}O{-}H} \\ \| \\ \mathrm{O} \end{array}$
Carbonic anhydride.........	$\mathbf{C}\mathrm{O}_2$.	O=C=O
Potassic carbonate	$\mathbf{C}\mathrm{OKo}_2$.	K—O—C—O—K

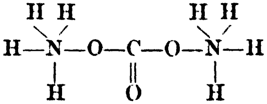
Symbolic. Graphic.

Marsh-gas............**CH_4**.

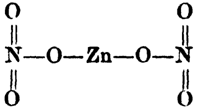


$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & — & C & — & H \\ & & | & & \\ & & H & & \end{array}$$

Ammonic carbonate, **$CO Amo_2$**.

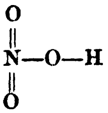


Zincic nitrate ...... $\left.\begin{array}{r} \mathbf{N}O_2O \\ Zn'' \\ \mathbf{N}O_2O \end{array}\right\}$.



$$\begin{array}{ccccccccc} O & & & & & & & & O \\ \| & & & & & & & & \| \\ N & — & O & — & Zn & — & O & — & N \\ \| & & & & & & & & \| \\ O & & & & & & & & O \end{array}$$

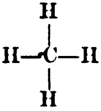
It must be carefully borne in mind that these graphic formulæ are intended to represent neither the shape of the molecules, nor the supposed relative position of the constituent hypothetical atoms. The lines connecting the different atoms of a compound, and which might with equal propriety be drawn in any other direction, provided they connected together the same elements, serve only to show the definite disposal of the bonds, the latter again being only a concrete symbolic expression of an abstract train of reasoning: thus the formula for nitric acid indicates that two of the three constituent atoms of oxygen are combined with nitrogen alone, and are consequently united to that element by both their bonds, whilst the third oxygen atom is combined both with nitrogen and hydrogen.



$$\begin{array}{ccccc} O & & & & \\ \| & & & & \\ N & — & O & — & H \\ \| & & & & \\ O & & & & \end{array}$$

The lines connecting the different atoms of a compound are but crude symbols of the bond of union between them; and it is scarcely necessary to remark that no such material connexions exist, the bonds which actually hold together the constituents of a compound being, as regards their nature, entirely unknown.

It deserves also to be here mentioned that graphic, like symbolic formulæ, are purely statical representations of chemical compounds; they take no cognizance of the amount of potential energy associated with the different elements. Thus in the formulæ for marsh-gas and carbonic anhydride,



$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

Marsh-gas.

$$O=C=O$$

Carbonic anhydride.

there is no indication that the molecule of the first compound contains a vast store of force, whilst the last is comparatively a powerless molecule.

CHAPTER IV.

COMPOUND RADICALS.

The term compound radical may be applied to any group of two or more atoms, which takes the place and performs the functions of an element in a chemical compound. In practice, however, it is only applied to any such group when the latter is met with in numerous chemical compounds.

An element is a *simple radical*, and enters into combination in the following manner (a, b, c, and d being monad elements, a'' a dyad, a''' a triad, and a^{iv} a tetrad element):—

$$\begin{array}{l} a + b = ab, \\ a'' + 2b = a''b_2, \\ a''' + 3b = a'''b_3, \\ \text{\&c.} \qquad \text{\&c.} \end{array}$$

A group of elements replacing a, a'', or a''' in the above equations is a *compound radical*, as in the following examples:—

$$\begin{aligned}
&(a''b) &&+\ b=(a''b)b,\\
&(a'''b)'' &&+2b=(a'''b)''b_2,\\
&(a'''bc) &&+\ b=(a'''bc)b,\\
&(a^{\mathrm{iv}}b)''' &&+3b=(a^{\mathrm{iv}}b)'''b_3,\\
&(a^{\mathrm{iv}}bc)'' &&+2b=(a^{\mathrm{iv}}bc)''b_2,\\
&(a^{\mathrm{iv}}bcd) &&+\ b=(a^{\mathrm{iv}}bcd)b.
\end{aligned}$$

The group of elements $(a''b)$ constitutes a compound monad radical equivalent to one atom of hydrogen or chlorine. The group $(a'''b)''$ is a compound dyad radical, &c. It is therefore evident that a polyad element is essential to every compound radical; in fact *a compound radical consists of one or more atoms of a polyad element in which one or more bonds are unsatisfied; and it is either a monad, dyad, triad, &c. radical, according to the number of monad atoms required to satisfy its active atomicity*. Such a radical, when a monad, triad, or pentad, cannot exist as a separate group; like hydrogen or nitrogen, when isolated, it combines with itself, forming a duplex molecule. It is only by the union of two atoms or groups of atoms that the vacated bonds can in these cases be satisfied.

From the above definition of a compound radical, it is evident that an almost infinite number of such bodies must exist; for in the compounds of every polyad element it is only necessary to vacate successive bonds to create each time a new compound radical. Thus marsh-gas CH_4 minus one atom of hydrogen gives the compound radical methyl CH_3; minus two atoms of hydrogen, it forms methylene $(CH_2)''$; and by the abstraction of three hydrogen atoms it is transformed into the triad radical formyl $(CH)'''$; but, except in a few cases, it is not advantageous thus to incorporate, as it were, compound radicals, which, instead of simplifying notation and nomenclature, would, if thus multiplied, only embarrass them. No compound radical, therefore, ought to receive recognition as such, unless it can be shown to enter into the composition of a large number of compounds.

The following are the names, symbols, and formulæ of the chief inorganic compound radicals recognized in the notation of this book:—

	Molecular formulæ.	Semimolecular formulæ.	Semimolecular symbols.
Hydroxyl	$(HO)_2$	HO	Ho.
Hydrosulphyl	$(HS)_2$	HS	Hs.
Ammonium	$(NH_4)_2$	NH_4	Am.
Ammonoxyl	$(NH_4O)_2$	NH_4O	Amo.
Amidogen	$(NH_2)_2$	NH_2	Ad.

In addition to these, certain compounds which metals form with oxygen are also regarded as compound radicals—for instance,

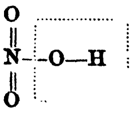
	Molecular formulæ.	Semimolecular or combining formulæ.	Semimolecular or combining symbols.
Potassoxyl.........	$(KO)_2$	KO	Ko.
Zincoxyl	(ZnO_2)	$\left\{\begin{matrix}O\\ Zn''\\ O\end{matrix}\right.$	Zno''.

The essential character of these last compound radicals is that the whole of the oxygen they contain is united with the metal by one bond only of each oxygen atom, as seen in the following graphic formulæ:—

Hydroxyl	$—O—H$
Potassoxyl	$—O—K$
Zincoxyl	$—O—Zn—O—$

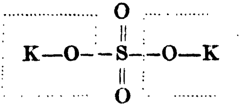
The metal thus becomes linked to other elements by these dyad atoms of oxygen. The functions of such compound radicals will be sufficiently evident from the following examples of compounds into which they enter, and in which their position is marked by dotted lines.

Nitric acid ....



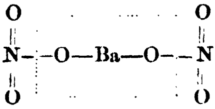
$$\begin{matrix}O & \\ \| & \\ N— & —O—H\\ \| & \\ O & \end{matrix}$$

Potassic sulphate ....



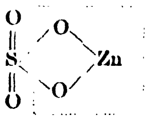
$$K{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}{-}O{-}K$$

Baric nitrate ....



$$\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{N}}{-}O{-}Ba{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{N}}$$

Zincic sulphate ....



It is not necessary to dignify all these metallic compound radicals with names; the chief point of importance about them is their abbreviated notation, in which the small letter o is attached to the symbol of the metal, the atomicity of the radical being marked in the usual manner. Although the small letter o in these symbols of combining quantities has no more reference to the composition of the radical than the d in the corresponding symbol of amidogen, yet it may usefully remind the reader that oxygen is always a constituent of the compound radicals so symbolized. It must be borne in mind that the number of atoms of oxygen in any radical of this class depends upon its atomicity: thus a monad contains only one atom of oxygen, a dyad two, and a triad always three atoms of oxygen. The use of any but monad and dyad metallic compound radicals is very rare.

CHAPTER V.

ATOMIC AND MOLECULAR COMBINATION.

In all the cases of chemical combination considered in the above Chapters, a union of atoms has been invariably contemplated. This atomic union is generally attended by the breaking up of previously existing molecules—two such molecules, by the mutual exchange of their atomic constituents, producing two new and perfectly distinct molecules. Thus when chlorine unites with hydrogen to form hydrochloric acid, a molecule of chlorine and one of hydrogen yield up their constituent atoms, forming two molecules of hydrochloric acid,

$$Cl_2 + H_2 = 2HCl.$$

In comparatively rare cases, two molecules combine to form only one new molecule; thus a molecule of carbonic oxide and one of chlorine combine to form one molecule of carbonic oxydichloride or phosgene gas: but the union is even here essentially atomic; for after combination both the oxygen and chlorine are directly united with the atom of carbon:

$$\underset{\text{Carbonic oxide.}}{\mathbf{C}''O} \quad + \quad \underset{\text{Chlorine.}}{Cl_2} \quad = \quad \underset{\text{Phosgene gas.}}{\mathbf{C}^{iv}OCl_2}.$$

Chemists are, however, compelled to admit an entirely different kind of union, which not unfrequently occurs, and which in conformity with the atomic hypothesis, may be appropriately termed ***molecular union*** or ***molecular combination***. In the formation of such compounds, no change takes place in the active atomicity of any of the molecules. It is this kind of combination which holds together salts and their water of crystallization, as, for instance,

Sodic chloride crystallized at $-10°$ C.......$NaCl, 2\mathbf{O}H_2$.
Sodic bromide crystallized below $+30°$ C...$NaBr, 2\mathbf{O}H_2$.
Sodic iodide crystallized below $+50°$ C...$NaI, 2\mathbf{O}H_2$.
Alum ....$\mathbf{S}_4O_8('Al'''_2O_6)^{vi}Ko_2, 24\mathbf{O}H_2$.

Numerous other instances of molecular combination might be adduced; but it is only necessary here to point out that such molecular unions will be distinguished from atomic combinations by the use of the comma, as in the above and following examples:

Tetramethylammonic tri-iodide ……	**N**Me_4 I, I_2.
Tetramethylammonic pentiodide ……	**N**Me_4 I, $2I_2$.
Tetramethylammonic iodo-dichloride	**N**Me_4 I, Cl_2.

In all cases molecular combination seems to be of a much more feeble character than atomic union; for, in the first place, such bodies are generally decomposed with facility; and secondly, the properties of their constituent molecules are markedly perceptible in the compounds. Thus the above so-called periodides of the organic bases greatly resemble iodine in appearance.

CHAPTER VI.

CLASSIFICATION OF ELEMENTS.

It has already been mentioned that the elements may be divided into two great classes, the metals and the non-metals or metalloids. A second division into positive and negative elements has also been explained. A third and still more important classification is founded upon the atomicity of the elements. In the following classified Table all three methods are embodied, the metalloids being printed in red type, and the metals in black, whilst the positive elements are printed in Roman characters, and the negative in italics. In addition, the different classes are also divided into sections, consisting of elements closely related in their chemical characters.

Monads.	Dyads.	Triads.	Tetrads.	Pentads.	Hexads.
1st Section. Hydrogen.	1st Section. *Oxygen.*	1st Section. Boron.	1st Section. Carbon. Silicon. Tin. Titanium.	1st Section. Nitrogen. Phosphorus. Vanadium. Arsenic. Antimony. Bismuth.	1st Section. Sulphur. Selenium. *Tellurium.*
2nd. Section. *Fluorine.* *Chlorine.* *Bromine.* *Iodine.*	2nd Section. Barium. Strontium. Calcium. Magnesium. Zinc.	2nd Section. Gold.	2nd Section. Thorium. Niobium. Tantalum. Zirconium. Aluminium. Indium.		2nd Section. Tungsten. Molybdenum.
3rd Section. Cæsium. Rubidium. Potassium. Sodium. Lithium.	3rd Section. Didymium. Lanthanum. Yttrium. Glucinum.	3rd Section. Thallium.	3rd Section. Platinum. Palladium.		3rd Section. Osmium. Iridium. Ruthenium. Rhodium.
4th Section. Silver.	4th Section. Cadmium. Mercury. Copper.		4th Section. Lead.		4th Section. Chromium. Manganese. Iron. Cobalt. Nickel. Uranium. Cerium.

CHAPTER VII.

WEIGHTS AND MEASURES.

THE weights and measures employed in this book are chiefly those of the French decimal system, founded upon the metre, which is $\frac{1}{10.000.000}$th part of a quadrant of a great terrestrial circle. The following Tables, published by Messrs. De La Rue and Co., will enable the student to convert these into their English equivalents whenever it may be necessary.

French Measures of Length.

	In English inches.	In English feet = 12 inches.	In English yards = 3 feet.	In English fathoms = 6 feet.	In English miles = 1760 yards.
Millimètre	0·03937	0·003281	0·0010936	0·0005468	0·0000006
Centimètre	0·39371	0·032809	0·0109363	0·0054682	0·0000062
Decimètre	3·93708	0·328090	0·1093633	0·0546816	0·0000621
Mètre	39·37079	3·280899	1·0936331	0·5468165	0·0006214
Décamètre	393·70790	32·808992	10·9363310	5·4681655	0·0062138
Hectomètre .	3937·07900	328·089920	109·3633100	54·6816550	0·0621382
Kilomètre	39370·79000	3280·899200	1093·6331000	546·8165500	0·6213824
Myriomètre ...	393707·90000	32808·992000	10936·3310000	5468·1655000	6·2138244

1 inch = 2·539954 centimètres.
1 foot = 3·0479449 decimètres.
1 yard = 0·9143835 mètre.
1 mile = 1·6093149 kilomètre.

French Measures of Surface.

	In English square feet.	In English square yards = 9 square feet.	In English poles = 272·25 sq. feet.	In English roods = 10890 sq. feet.	In English acres = 43560 sq. feet.
Centiare or sq. mètre	10·7642993	1·1960333	0·0395383	0·0009885	0·0002471
Are or 100 sq. mètres	1076·4299342	119·6033260	3·9538290	0·0988457	0·0247114
Hectare or 10,000 square mètres	107642·993418	11960·332602	395·382950	9·8845724	2·4711431

1 square inch = 6·4513669 square centimètres.
1 square foot = 9·2899683 square decimètres.
1 square yard = 0·8360715 square mètre or centiare.
1 acre = 0·4046710 hectare.

French Measures of Capacity.

	In cubic inches.	In cubic feet = 1728 cubic inches.	In pints = 34·65923 cubic inches.	In gallons = 8 pints = 277·27384 cubic inches.	In bushels = 8 gallons = 2218·19075 cubic inches.
Millilitre or cubic centimètre	0·06103	0·000035	0·00176	0·0002201	0·0000275
Centilitre or 10 cubic centimètres	0·61027	0·000353	0·01761	0·0022010	0·0002751
Décilitre or 100 cubic centimètres	6·10271	0·003532	0·17608	0·0220097	0·0027512
Litre or cubic décimètre	61·02705	0·035317	1·76077	0·2200967	0·0275121
Décalitre or centistère	610·27052	0·353166	17·60773	2·2009668	0·2751208
Hectolitre or décistère	6102·70515	3·531658	176·07734	22·0096677	2·7512085
Kilolitre, or Stère, or cubic mètre	61027·05152	35·316581	1760·77341	220·0966767	27·5120846
Myriolitre or décastère	610270·51519	353·165807	17607·73414	2200·9667675	275·1208459

1 cubic inch = 16·386176 cubic centimètres.
1 cubic foot = 28·315312 cubic décimètres, or litres.
1 gallon = 4·543458 litres.

French Measures of Weight.

	In English grains.	In Troy ounces= 480 grains.	In avoirdupois lbs.= 7000 grains.	In cwts.= 112 lbs.= 784000 grs.	Ton= 20 cwts.= 15680000 grs.
Milligramme	0·01543	0·000032	0·0000022	0·0000000	0·0000000
Centigramme	0·15432	0·000322	0·0000220	0·0000002	0·0000000
Décigramme	1·54323	0·003215	0·0002205	0·0000020	0·0000001
Gramme	15·43235	0·032151	0·0022046	0·0000197	0·0000010
Décagramme	154·32349	0·321507	0·0220462	0·0001968	0·0000098
Hectogramme	1543·23488	3·215073	0·2204621	0·0019684	0·0000984
Kilogramme	15432·34880	32·150727	2·2046213	0·0196841	0·0009842
Myriogramme	154323·48800	321·507267	22·0462126	0·1968412	0·0298421

1 grain = 0·064792 gramme. 1 lb. avoir. = 0·453593 kilogr.
1 troy oz. = 31·103496 grammes. 1 cwt. = 50·802377 kilogrs.
1 metrekilogramme = 7·2324 foot lbs.

Temperatures are expressed upon the Centigrade scale; and barometric measurements are given in millimetres.

For the ready conversion of gaseous volumes into weights, I have adopted the *crith*, or standard multiple proposed by Dr. Hofmann. The crith is the weight of one litre, or cubic decimetre, of hydrogen at 0° C. and at a pressure of 760 millimetres of mercury. The following is Dr. Hofmann's description of the value and applications of this unit.

"The actual weight of this cube of hydrogen, at the standard temperature and pressure mentioned, is 0·0896 gramme; a figure which I earnestly beg you to inscribe, as with a sharp graving tool, upon your memory. There is probably no figure in chemical science more important than this one to be borne in mind, and to be kept ever in readiness for use in calculation at a moment's notice. For this litre-weight of hydrogen =0·0896 gramme (I purposely repeat it) is the standard multiple, or coefficient, by means of which the weight of one litre of any other gas, simple or compound, is computed. Again, therefore, I say, do not let slip this figure—0·0896 gramme. So important, indeed, is this standard weight unit, that some name—the simpler and briefer the better—is needed to denote it. For this purpose I venture to suggest the term *crith*, derived from the Greek word κριθὴ, signifying a barley-corn, and figuratively employed to imply a small weight. The weight of 1 litre of hydrogen being called 1 crith, the volume-weight of other gases, referred to hydrogen as a standard, may be expressed in terms of this unit.

"For example, the relative volume-weight of chlorine being 35·5, that of oxygen 16, that of nitrogen 14, the actual weight of 1 litre of each of these elementary gases, at 0° C. and $0^{m}\cdot76$ pressure, may be called respectively 35·5 *criths*, 16 *criths*, and 14 *criths*.

"So, again, with reference to the compound gases; the relative volume-weight of each is equal to half the weight of its product-volume. Hydrochloric acid (HCl), for example, consists of 1 vol. of hydrogen + 1 vol. of chlorine = 2 volumes; or, by weight, $1+35\cdot5=36\cdot5$ units; whence it follows that the relative volume-weight of hydrochloric acid gas is $\frac{36\cdot5}{2}=18\cdot25$ units; which last figure therefore expresses the number of *criths* which one litre of hydrochloric acid gas weighs at 0° C. temperature and $0^{m}\cdot76$ pressure; and the crith being (as I trust you already bear in mind) 0·0896 gramme, we have

$$18\cdot25\times0\cdot0896=1\cdot6352$$

as the actual weight in grammes of hydrochloric acid gas.

"So, once more, as the product-volume of water-gas (H_2O) (taken at the above temperature and pressure) contains 2 vols. of hydrogen + 1 vol. of oxygen, and therefore weighs $2+16=18$ units, the single volume of water-gas weighs $\frac{18}{2}=9$ units; or, substituting as before the concrete for the abstract value, 1 litre of water-gas weighs 9 *criths*; that is to say, $9\times0\cdot0896$ gramme, $=0\cdot8064$ gramme.

"In like manner the product-volume of sulphuretted hydrogen (H_2S) = 2 litres of hydrogen, weighing 2 criths, + 1 litre of sulphur-gas, weighing 32 criths, together $2+32=34$ criths, which, divided by 2, gives $\frac{34}{2}=17$ criths $=17\times0\cdot0896$ gramme $=1\cdot5232$ gramme = the weight of 1 litre of sulphuretted hydrogen at standard temperature and pressure.

"And so, lastly, of ammonia (NH_3); it contains in 2 litres 3 litres of hydrogen, weighing 3 criths, and 1 litre of nitrogen, weighing 14 criths; its total product-volume-weight is therefore 3 + 14 = 17 criths, and its single volume or litre weight is consequently

$$\tfrac{17}{2} = 8·5 \text{ criths} = 8·5 \times 0·0896 \text{ gramme} = 0·7616 \text{ gramme.}$$

"Thus, by the aid of the hydrogen-litre-weight or *crith* = 0·0896 gramme, employed as a common multiple, the actual or concrete weight of 1 litre of any gas, simple or compound, at standard temperature and pressure, may be deduced from the mere abstract figure expressing its volume-weight relatively to hydrogen."

The number expressing in criths the weight of 1 litre of any gas or vapour being identical with its specific gravity compared with hydrogen taken as unity, it is easy, when this number is known, to calculate the specific gravity of the gas compared with air taken as unity. For this purpose it is only necessary to multiply by ·0693, which is the specific gravity of hydrogen compared with air = 1.

Thus the specific gravity of oxygen compared with air is

$$16 \times ·0693 = 1·1088\,;$$

of chlorine,

$$35·5 \times ·0693 = 2·46015\,;$$

of hydrochloric acid,

$$18·25 \times ·0693 = 1·264725.$$

CHAPTER VIII.

MONAD ELEMENTS.

SECTION I.

HYDROGEN, H_2.

Atomic weight = 1. *Molecular weight* = 2. *Molecular volume* ☐☐. 1 *litre weighs* 1 *crith. Atomicity ', being the standard of comparison.*

Occurrence.—In combination, as water, in very large quantities in nature. In almost all vegetable and animal substances, and in many minerals. In the free state in the gases of volcanoes. In the sun, certain stars, and nebulæ.

Preparation.—1. By the action of sodium upon water:—

$$\underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Sodium.}}{Na_2} = \underset{\text{Sodic hydrate.}}{2\mathbf{O}NaH} + \underset{\text{Hydrogen.}}{H_2}.$$

2. By the action of sodium upon dry hydrochloric acid:—

$$\underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Sodium.}}{Na_2} = \underset{\text{Sodic chloride.}}{2NaCl} + \underset{\text{Hydrogen.}}{H_2}.$$

3. By the action of zinc, iron, or certain other metals on hydrochloric acid:—

$$\underset{\text{Hydrochloric acid.}}{2HCl} + Zn = \underset{\text{Zincic chloride.}}{\mathbf{Zn}Cl_2} + H_2.$$

4. By the action of zinc, or certain other metals on dilute sulphuric acid:—

$$\underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + Zn = \underset{\text{Zincic sulphate.}}{\mathbf{S}O_2Zno''} + H_2.$$

5. By passing steam over iron heated to redness:—

$$Fe_3 + \underset{\text{Water.}}{4\mathbf{O}H_2} = \underset{\text{Triferric tetroxide.}}{{}^{iv}(\mathbf{Fe}_3)^{viii}O_4} + 4H_2.$$

6. By the action of zinc on a boiling solution of potassic hydrate:—

$$\underset{\text{Potassic hydrate.}}{2\mathbf{O}KH} + Zn = \underset{\text{Potassic zinc oxide.}}{\mathbf{Zn}Ko_2} + H_2.$$

7. By the electrolysis of water and of some other liquids containing hydrogen.

8. By the action of intense heat upon water.

9. In the destructive distillation of some organic substances.

SECTION II.

CHLORINE, Cl_2.

Atomic weight=35·5. *Molecular weight*=71. *Molecular volume* ☐☐. 1 *litre weighs* 35·5 *criths. Has not been solidified. Liquefies at* 15°·5 C., *under a pressure of* 4 *atmospheres. Atomicity* $'$. *Evidence of atomicity*, HCl.

Occurrence.—Always in combination—with sodium and other metals in sea-water, and in the solid state in the salt-beds of Cheshire, Worcester, &c. Evolved from volcanoes in the form of hydrochloric acid.

Preparation.—1. By heating certain metallic chlorides, as platinic and auric chlorides:—

$$\underset{\text{Platinic chloride.}}{\mathbf{Pt}Cl_4} = 2Cl_2 + Pt.$$

2. By gently heating a mixture of manganic oxide and hydrochloric acid, when the reaction takes place in two stages:—

$$\underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Manganic chloride.}}{\mathbf{Mn}Cl_4} + \underset{\text{Watr.}}{2\mathbf{O}H_2};$$

$$\underset{\text{Manganic chloride.}}{\mathbf{Mn}Cl_4} = \underset{\text{Manganous chloride.}}{\mathbf{Mn}Cl_2} + Cl_2.$$

3. By heating a mixture of sulphuric acid, sodic chloride, and manganic oxide, when the whole of the chlorine present is liberated:—

$$\underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} + \underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} + \underset{\text{Sodic chloride.}}{2NaCl} = \underset{\text{Sodic sulphate.}}{\mathbf{S}O_2Nao_2} +$$

$$\underset{\text{Manganous sulphate.}}{\mathbf{S}O_2Mno''} + \underset{\text{Water.}}{2\mathbf{O}H_2} + Cl_2.$$

If in the second process a mixture of manganic oxide, hydrochloric acid, and sulphuric acid be employed, the whole of the chlorine is evolved:—

$$\underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + \underset{\text{Hydrochloric acid.}}{2HCl} =$$

$$\underset{\text{Manganous sulphate.}}{\mathbf{S}O_2Mno''} + \underset{\text{Water.}}{2\mathbf{O}H_2} + Cl_2.$$

4. By the electrolysis of hydrochloric acid.

5. Cuprous chloride, heated in air and moistened with hydrochloric acid or, heated in a current of hydrochloric acid gas, steam and air, produces cupric chloride. By raising the temperature chlorine is evolved, and cuprous chloride reproduced:—

$$\mathbf{Cu}_2Cl_2 \ + \ O \ + \ 2HCl \ = \ 2\mathbf{Cu}Cl_2 \ + \ \mathbf{O}H_2,$$

and

$$2\mathbf{Cu}Cl_2 \ = \ \mathbf{Cu}_2Cl_2 \ + \ Cl_2.$$

Reactions.—1. Chlorine and hydrogen, when mixed, unite instantly, with explosion, under the influence of sunlight, or of powerful artificial light, or on the application of a burning body to the mixture. An igniting jet of hydrogen continues to burn when plunged into chlorine:

$$H_2 \ + \ Cl_2 \ = \ 2HCl.$$

2. Chlorine has so great an attraction for hydrogen, that it removes the latter from its compounds with carbon. When a rag moistened with turpentine is plunged into chlorine, the chlorine and hydrogen unite, with evolution of heat and light, carbon being liberated:—

$$\underset{\text{Turpentine.}}{C_{10}H_{16}} \ + \ 8Cl_2 \ = \ \underset{\text{Hydrochloric acid.}}{16HCl} \ + \ 10C.$$

HYDROCHLORIC ACID, *Chlorhydric Acid, Muriatic Acid.*

HCl.

Molecular weight =36·5. *Molecular volume* □□. 1 *litre weighs* 18·25 *criths. Has not been solidified. Condenses at* 10°, *under a pressure of* 40 *atmospheres.*

Occurrence.—Evolved from volcanoes.

Preparation.—1. From its elements, as above described.

2. By gently heating sodic chloride with sulphuric acid, previously diluted with a small quantity of water:—

$$\mathbf{S}O_2Ho_2 + NaCl = \mathbf{S}O_2HoNao + HCl;$$

Sulphuric acid. Sodic chloride. Hydric sodic sulphate. Hydrochloric acid.

or $$\mathbf{S}O_2Ho_2 + 2NaCl = \mathbf{S}O_2Nao_2 + 2HCl.$$

Sulphuric acid. Sodic chloride. Sodic sulphate. Hydrochloric acid.

Reactions.—Hydrochloric acid may be converted into salts, termed chlorides, by the action of certain metals, as described above, and also by that of the metallic hydrates, or oxides:—

$$\mathbf{O}KH + HCl = KCl + \mathbf{O}H_2.$$

Potassic hydrate. Hydrochloric acid. Potassic chloride. Water.

$$\mathbf{Zn}O + 2HCl = \mathbf{Zn}Cl_2 + \mathbf{O}H_2.$$

Zincic oxide. Hydrochloric acid. Zincic chloride. Water.

For the remaining monad elements of this Section, see Chapter XIV.

CHAPTER IX.

DYAD ELEMENTS.

SECTION I.

OXYGEN, O_2.

Atomic weight=16. *Molecular weight*=32. *Molecular volume* ☐☐. 1 *litre weighs* 16 *criths*. *Atomicity* $''$. *Evidence of atomicity*:—

Water	$\mathbf{O}H_2$.
Potassic hydrate	$\mathbf{O}KH$.
Argentic oxide	$\mathbf{O}Ag_2$.
Hypochlorous anhydride	$\mathbf{O}Cl_2$.

Occurrence.—In the free state, in the atmosphere. In the combined state, in water, in most mineral bodies, and in almost all animal and vegetable compounds.

Preparation.—1. If metallic mercury be heated to its boiling-point with access of air, it gradually absorbs oxygen, being converted into mercuric oxide, $\mathbf{Hg}''O$. This compound, when more strongly heated, is resolved into its elements,

$$\underset{\text{Mercuric oxide.}}{2\mathbf{Hg}O} = \underset{\text{Mercury.}}{2Hg} + \underset{\text{Oxygen.}}{O_2}.$$

2. By heating native manganic oxide (pyrolusite), a portion of its oxygen is liberated:—

$$\underset{\text{Manganic oxide.}}{3\mathbf{Mn}O_2} = \underset{\text{Trimanganic tetroxide.}}{{}^{iv}(\mathbf{Mn}_3)^{viii}O_4} + O_2.$$

3. Oxygen is evolved in nature, in a remarkable manner, by the decomposition of carbonic anhydride, $\mathbf{C}O_2$, by the green leaves of plants, the vegetable assimilating the carbon, whilst the oxygen escapes into the atmosphere:—

$$\underset{\text{Carbonic anhydride.}}{\mathbf{C}O_2} = C + O_2.$$

4. By the action of heat upon potassic chlorate:—

$$\text{(Atomic)} \quad \left\{\begin{matrix}\mathbf{O}Cl\\ O\\ \mathbf{O}K\end{matrix}\right. = KCl + 3O;\ \text{or}$$

$$\text{(Molecular)} \quad \underset{\text{Potassic chlorate.}}{2\left\{\begin{matrix}\mathbf{O}Cl\\ O\\ \mathbf{O}K\end{matrix}\right.} = \underset{\text{Potassic chloride.}}{2KCl} + \underset{\text{Oxygen.}}{3O_2}.$$

5. By mixing the potassic chlorate with manganic oxide, the oxygen is evolved at a much lower temperature; the manganic oxide appears to take no part in the reaction.

6. By dropping concentrated sulphuric acid into a red-hot platinum retort, the acid is decomposed into oxygen, sulphurous anhydride, and water:—

$$\text{(Atomic)} \quad \mathbf{S}O_2Ho_2 = \mathbf{S}O_2 + \mathbf{O}H_2 + O;\ \text{or}$$

$$\text{(Molecular)} \quad \underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} = \underset{\text{Sulphurous anhydride.}}{2\mathbf{S}O_2} + \underset{\text{Water.}}{2\mathbf{O}H_2} + O_2.$$

7. By the electrolysis of dilute sulphuric acid.

8. By the action of heat upon a mixture of manganic oxide and sulphuric acid :—

$$(\text{Atomic})\quad \mathbf{Mn}O_2 + \mathbf{S}O_2Ho_2 = \mathbf{S}O_2Mno'' + \mathbf{O}H_2 + O;\ \text{or}$$

$$(\text{Molecular})\, \underset{\text{Manganic oxide.}}{2\mathbf{Mn}O_2} + \underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} = \underset{\text{Manganous sulphate.}}{2\mathbf{S}O_2Mno''} + \underset{\text{Water.}}{2\mathbf{O}H_2} + O_2.$$

9. By heating a mixture of potassic bichromate and sulphuric acid :—

$$\underset{\text{Potassic bichromate.}}{2\left\{\begin{matrix}\mathbf{Cr}O_2Ko \\ O \\ \mathbf{Cr}O_2Ko\end{matrix}\right.} + \underset{\text{Sulphuric acid.}}{8\mathbf{S}O_2Ho_2} = \underset{\text{Potassic sulphate.}}{2\mathbf{S}O_2Ko_2}$$

$$+ \underset{\text{Chromic sulphate.}}{2\mathbf{S}_3O_6('Cr'''_2O_6)^{vi}} + 3O_2 + \underset{\text{Water.}}{8\mathbf{O}H_2}.$$

10. By passing steam and chlorine through a red-hot porcelain tube, hydrochloric acid and oxygen are formed :—

$$\underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Chlorine.}}{2Cl_2} = \underset{\text{Hydrochloric acid.}}{4HCl} + O_2.$$

Reaction.—A mixture of two volumes of hydrogen and one volume of oxygen explodes at a red heat, water being produced. The same compound is formed when hydrogen is burnt in oxygen, or oxygen in hydrogen :—

$$(\text{Molecular})\quad 2H_2 + O_2 = 2\mathbf{O}H_2.$$

ALLOTROPIC OXYGEN, or OZONE, O_3.

Molecular weight=48. *Molecular volume* ☐☐. 1 *litre weighs* 24 *criths.*

Preparation.—1. When electric sparks are passed through air, or oxygen, a peculiar odour, which is due to ozone, is observed.

2. By placing phosphorus for a few hours in moist air, at about the ordinary temperature.

3. By passing an electric current through dilute sulphuric or chromic acid.

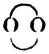
Thus obtained, ozone is always mixed with a large proportion of air, or oxygen.

Properties.—Powerfully oxidizing. It oxidizes organic matter, silver, and mercury at the ordinary temperature. When oxygen is converted into ozone, contraction of volume takes place; and when the ozone is heated to 290°, it is retransformed into the original volume of ordinary oxygen,—indicating that the molecule of ozone contains more atoms than the molecule of ordinary oxygen.

In most cases of oxidation by ozone no diminution of the volume of gas takes place, the additional atoms, previously introduced into the molecules of oxygen, being removed, and ordinary oxygen becoming free. But oil of turpentine absorbs the whole molecule of the ozone, leaving untouched the oxygen which was previously present in the state of admixture. By observing the contraction during its production, and the diminution of volume produced by absorbing the ozone with oil of turpentine, its density may be readily calculated, and consequently its atomic constitution.

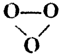
By this means the specific gravity of ozone has been shown to be 24, the molecular weight being therefore 48, which is the weight of 3 atoms of oxygen.

In ordinary oxygen, the molecule is composed of two atoms of oxygen, and is represented by



weighing 32.

In ozone, the molecule contains 3 atoms of oxygen, and is represented by



weighing 48.

WATER, *Hydric Oxide.*

$H—O—H \qquad OH_2.$

Molecular weight $=18$. ***Molecular volume*** $\square\square$. 1 *litre of water-vapour weighs* 9 *criths.* ***Melts at*** 0°. ***Boils at*** 100°.

Occurrence.—Very abundantly in nature.

Formation.—1. By the direct union of hydrogen and oxygen, as already described.

2. As a secondary product, in numberless chemical reactions, as, for instance, in the action of hydrochloric acid on potassic hydrate:—

$$\underset{\text{Potassic hydrate.}}{\mathbf{O}KH} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Water.}}{\mathbf{O}H_2} + \underset{\text{Potassic chloride.}}{KCl.}$$

Reactions.—1. By its action many metallic oxides are converted into hydrates:—

$$\underset{\text{Potassic oxide.}}{\mathbf{O}K_2} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Potassic hydrate.}}{2\mathbf{O}KH.}$$

$$\underset{\text{Baric oxide.}}{\mathbf{Ba}O} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Baric hydrate.}}{\mathbf{Ba}Ho_2.}$$

2. It transforms anhydrides into acids:—

$$\underset{\text{Nitric anhydride.}}{\mathbf{N}_2O_5} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho.}$$

$$\underset{\text{Sulphuric anhydride.}}{\mathbf{S}O_3} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2.}$$

$$\underset{\text{Phosphoric anhydride.}}{\mathbf{P}_2O_5} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\text{Phosphoric acid.}}{2\mathbf{P}OHo_3.}$$

3. It also unites molecularly with many compounds, as water of crystallization (see Chapter V.), as in the following instances:—

$\mathbf{Ba}Cl_2, 2\mathbf{O}H_2$	Baric chloride.
$\mathbf{S}O_2Nao_2, 10\mathbf{O}H_2$	Sodic sulphate.
$\mathbf{S}_4O_8Ko_2('Al'''_2O_6)^{vi}, 24\mathbf{O}H_2$	Potash alum.

HYDROXYL, *Hydric Peroxide.*

$$\mathrm{H{-}O{-}\vdots{-}O{-}H} \qquad \mathrm{H_2O_2} \text{ or } \mathrm{(HO)_2} \text{ or } \mathrm{Ho_2} \text{ or } \left\{\begin{matrix}\mathbf{O}\mathrm{H}\\ \mathbf{O}\mathrm{H}\end{matrix}\right. .$$

Probable molecular weight = 34.

Preparation.—By passing a current of carbonic anhydride through water in which baric peroxide is suspended:—

$$\underset{\text{Baric peroxide.}}{\left\{\begin{matrix}\mathbf{O}\\ \mathbf{O}\end{matrix}\right.\mathrm{Ba''}} + \underset{\text{Carbonic anhydride.}}{\mathbf{C}\mathrm{O_2}} + \underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}} = \underset{\text{Baric carbonate.}}{\mathbf{C}\mathrm{OBao''}} + \underset{\text{Hydroxyl.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{H}\\ \mathbf{O}\mathrm{H}\end{matrix}\right.}$$

Reactions.—1. By heat it is decomposed into water and oxygen:—

$$\underset{\text{Hydroxyl.}}{2\left\{\begin{matrix}\mathbf{O}\mathrm{H}\\ \mathbf{O}\mathrm{H}\end{matrix}\right.} = \underset{\text{Water.}}{2\mathbf{O}\mathrm{H_2}} + \underset{\text{Oxygen.}}{\mathrm{O_2}}.$$

2. Hydroxyl is transformed into water by the action of nascent hydrogen: if hydroxyl be introduced into an apparatus in which hydrogen is being generated, the gas ceases to be evolved:—

$$\underset{\text{Sulphuric acid.}}{\mathbf{S}\mathrm{O_2Ho_2}} + \mathrm{Zn} + \underset{\text{Hydroxyl.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{H}\\ \mathbf{O}\mathrm{H}\end{matrix}\right.} = \underset{\text{Zincic sulphate.}}{\mathbf{S}\mathrm{O_2Zno''}} + \underset{\text{Water.}}{2\mathbf{O}\mathrm{H_2}}.$$

3. Hydroxyl liberates iodine from potassic iodide:—

$$\underset{\text{Potassic iodide.}}{2\mathrm{KI}} + \underset{\text{Hydroxyl.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{H}\\ \mathbf{O}\mathrm{H}\end{matrix}\right.} = \underset{\text{Potassic hydrate.}}{2\mathbf{O}\mathrm{KH}} + \underset{\text{Iodine.}}{\mathrm{I_2}}.$$

4. Hydroxyl is a powerful oxidizing agent; it converts, for instance, plumbic sulphide into plumbic sulphate:—

$$\underset{\text{Plumbic sulphide.}}{\mathbf{Pb}\mathrm{S''}} + \underset{\text{Hydroxyl.}}{4\left\{\begin{matrix}\mathbf{O}\mathrm{H}\\ \mathbf{O}\mathrm{H}\end{matrix}\right.} = \underset{\text{Plumbic sulphate.}}{\mathbf{S}\mathrm{O_2Pbo''}} + \underset{\text{Water.}}{4\mathbf{O}\mathrm{H_2}}.$$

COMPOUNDS OF CHLORINE WITH OXYGEN AND HYDROXYL.

Oxygen forms many compounds with chlorine and with chlorine and hydroxyl; but none of them can be produced by direct combination. The following list contains all that are known:—

Hypochlorous anhydride	$\mathbf{O}Cl_2$.	Cl—O—Cl
Chloric oxide? ...	$\begin{cases}\mathbf{O}Cl\\ \mathbf{O}Cl\end{cases}$.	Cl—O—O—Cl
Chlorous anhydride	$\begin{cases}\mathbf{O}Cl\\ O\\ \mathbf{O}Cl\end{cases}$.	Cl—O—O—O—Cl
Chloric peroxide	$\begin{cases}\mathbf{O}Cl\\ O\\ O\\ \mathbf{O}Cl\end{cases}$.	Cl—O—O—O—O—Cl
Chloric hyper-oxide?	$\begin{cases}\mathbf{O}Cl\\ O\\ O\\ O\\ O\\ \mathbf{O}Cl\end{cases}$.	Cl—O—O—O—O—O—O—Cl
Hypochlorous acid,	$\mathbf{O}ClH$, or $ClHo$.	H—O—Cl
Chlorous acid...	$\mathbf{O}ClHo$, or $\begin{cases}\mathbf{O}Cl\\ \mathbf{O}H\end{cases}$.	H—O—O—Cl
Chloric acid...	$\begin{cases}\mathbf{O}Cl\\ \mathbf{O}Ho\end{cases}$, or $\begin{cases}\mathbf{O}Cl\\ O\\ \mathbf{O}H\end{cases}$.	H—O—O—O—Cl
Perchloric acid	$\begin{cases}\mathbf{O}Cl\\ O\\ \mathbf{O}Ho\end{cases}$, or $\begin{cases}\mathbf{O}Cl\\ O\\ O\\ \mathbf{O}H\end{cases}$.	H—O—O—O—O—Cl

HYPOCHLOROUS ANHYDRIDE.

$\mathbf{O}Cl_2$.

Molecular weight =87. *Molecular volume* ☐☐. 1 *litre of hypochlorous anhydride vapour weighs* 43·5 *criths. Boils at about* 20°.

Preparation.—By passing chlorine over dry mercuric oxide, at a low temperature:—

$$\underset{\text{Mercuric oxide.}}{2\mathbf{Hg}O} + 2Cl_2 = \underset{\text{Mercuric oxychloride.}}{\left\{\begin{array}{l}\mathbf{Hg}Cl\\ O\\ \mathbf{Hg}Cl\end{array}\right.} + \underset{\text{Hypochlorous anhydride.}}{\mathbf{O}Cl_2}.$$

CHLOROUS ANHYDRIDE.

$$\left\{\begin{array}{l}\mathbf{O}Cl\\ O\\ \mathbf{O}Cl\end{array}\right.$$

Molecular weight =119. *Molecular volume anomalous* ⌈☐☐⌉. 1 *litre weighs* 39·7 *criths.*

Preparation.—By gently heating in a water-bath a mixture of potassic chlorate, nitric acid, and arsenious acid. Four different reactions are to be distinguished in this operation:—

$$1.\quad \underset{\text{Potassic chlorate}}{\left\{\begin{array}{l}\mathbf{O}Cl\\ \mathbf{O}Ko\end{array}\right.} + \underset{\text{Nitric acid}}{\mathbf{N}O_2Ho} = \underset{\text{Chloric acid}}{\left\{\begin{array}{l}\mathbf{O}Cl\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko};$$

$$2.\quad \underset{\text{Arsenious acid}}{\mathbf{As}Ho_3} + \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho} = \underset{\text{Nitrous acid}}{\mathbf{N}OHo} + \underset{\text{Arsenic acid}}{\mathbf{As}OHo_3};$$

$$3.\quad \underset{\text{Chloric acid.}}{\left\{\begin{array}{l}\mathbf{O}Cl\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Nitrous acid.}}{\mathbf{N}OHo} = \underset{\text{Chlorous acid}}{\mathbf{O}ClHo} + \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho};$$

$$4.\quad \underset{\text{Chlorous acid.}}{2\mathbf{O}ClHo} = \underset{\text{Chlorous anhydride}}{\left\{\begin{array}{l}\mathbf{O}Cl\\ O\\ \mathbf{O}Cl\end{array}\right.} + \underset{\text{Water}}{\mathbf{O}H_2}.$$

CHLORIC PEROXIDE.

$$\left\{\begin{matrix}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Cl}\end{matrix}\right.$$

Molecular weight =135. *Boils at* 20°.

Preparation.—By the action of sulphuric acid on potassic chlorate:—

$$\underset{\text{Potassic chlorate.}}{3\left\{\begin{matrix}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right.}+\underset{\text{Sulphuric acid.}}{2\mathbf{S}\mathrm{O_2Ho_2}}=\underset{\text{Potassic perchlorate.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right.}+\underset{\text{Hydric potassic sulphate.}}{2\mathbf{S}\mathrm{O_2HoKo}}+\underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}}+\underset{\text{Chloric peroxide}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Cl}\end{matrix}\right.}$$

HYPOCHLOROUS ACID.

$\mathbf{O}\mathrm{ClH}$, or ClHo.

Molecular weight =52·5.

Preparation.—1. By the action of water on hypochlorous anhydride:—

$$\underset{\text{Hypochlorous anhydride.}}{\mathbf{O}\mathrm{Cl_2}}\quad+\quad\underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}}\quad=\quad\underset{\text{Hypochlorous acid.}}{2\mathrm{ClHo}.}$$

2. By the action of chlorine upon mercuric oxide, in the presence of water:—

$$\underset{\text{Mercuric oxide.}}{2\mathbf{Hg}\mathrm{O}}\quad+\quad\underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}}\quad+\quad\underset{\text{Chlorine.}}{2\mathrm{Cl_2}}\quad=\quad\underset{\text{Mercuric oxychloride.}}{\left\{\begin{matrix}\mathbf{Hg}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{Hg}\mathrm{Cl}\end{matrix}\right.}\quad+\quad\underset{\text{Hypochlorous acid.}}{2\mathrm{ClHo}.}$$

Reactions.—1. By the action of hydrochloric acid, chlorine is evolved from both the hydrochloric acid and hypochlorous acid:—

$$\underset{\text{Hypochlorous acid.}}{\mathrm{ClHo}}\quad+\quad\underset{\text{Hydrochloric acid.}}{\mathrm{HCl}}\quad=\quad\underset{\text{Chlorine.}}{\mathrm{Cl_2}}\quad+\quad\underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}.}$$

2. By the action of argentic oxide, oxygen is evolved from both compounds :—

$$\underset{\text{Argentic oxide.}}{\mathbf{O}Ag_2} + \underset{\text{Hypochlorous acid.}}{2ClHo} = \underset{\text{Argentic chloride.}}{2AgCl} + \underset{\text{Water.}}{\mathbf{O}H_2} + O_2.$$

3. By the action of hypochlorous acid, metallic oxides or hydrates are converted into hypochlorites :—

$$\underset{\text{Potassic hydrate.}}{\mathbf{O}KH} + \underset{\text{Hypochlorous acid.}}{ClHo} = \underset{\text{Potassic hypochlorite.}}{ClKo} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

It was formerly supposed that hypochlorites, together with chlorides, were formed when chlorine acted upon certain metallic oxides and hydrates :—

$$2Cl_2 + \underset{\text{Calcic hydrate.}}{2\mathbf{Ca}Ho_2} = \underset{\text{Calcic chloride.}}{\mathbf{Ca}Cl_2} + \underset{\text{Calcic hypochlorite.}}{Cao''Cl_2} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

But the so-called chloride of lime or bleaching-powder does not contain calcic chloride, and the true reaction appears to be

$$\underset{\text{Calcic hydrate.}}{\mathbf{Ca}Ho_2} + Cl_2 = \underset{\text{Bleaching-powder*.}}{\mathbf{Ca}(OCl)Cl} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

By the action of acids this compound yields free chlorine :—

$$\underset{\text{Bleaching-powder.}}{\mathbf{Ca}(OCl)Cl} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Calcic sulphate.}}{\mathbf{S}O_2Cao''} + \underset{\text{Water.}}{\mathbf{O}H_2} + Cl_2.$$

CHLOROUS ACID.

$$\mathbf{O}ClHo \text{ or } \left\{\begin{matrix}\mathbf{O}Cl \\ \mathbf{O}H\end{matrix}\right. .$$

Molecular weight =68·5.

Preparation.—By the action of water upon chlorous anhydride :—

$$\underset{\text{Chlorous anhydride.}}{\left\{\begin{matrix}\mathbf{O}Cl \\ \mathbf{O} \\ \mathbf{O}Cl\end{matrix}\right.} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Chlorous acid.}}{2\left\{\begin{matrix}\mathbf{O}Cl \\ \mathbf{O}H.\end{matrix}\right.}$$

---

* Cl—Ca—O—Cl.

CHLORIC ACID.

$$\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ho}\end{array}\right. \quad \text{or} \quad \left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{H}\end{array}\right. .$$

Molecular weight =84·5.

Preparation.—By the action of dilute sulphuric acid upon baric chlorate:—

$$\underset{\text{chlorate.}}{\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathrm{Bao}''\\ \mathrm{O}\\ \mathbf{O}\mathrm{Cl}\end{array}\right.} \quad + \quad \underset{\text{Sulphuric acid.}}{\mathbf{S}\mathrm{O}_2\mathrm{Ho}_2} \quad = \quad \underset{\text{Chloric acid.}}{2\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ho}\end{array}\right.} \quad + \quad \underset{\text{Baric sulphate.}}{\mathbf{S}\mathrm{O}_2\mathrm{Bao}''}.$$

Decomposition.—By boiling, it is decomposed into perchloric acid, water, chlorine, and oxygen:—

$$\underset{\text{Chloric acid.}}{3\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ho}\end{array}\right.} \quad = \quad \underset{\text{Perchloric acid.}}{\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ho}\end{array}\right.} \quad + \quad \underset{\text{Water.}}{\mathbf{O}\mathrm{H}_2} \quad + \quad \mathrm{Cl}_2 \quad + \quad 2\mathrm{O}_2.$$

Preparation of Chlorates.—1. Potassic chlorate may be prepared by the action of chlorine upon a concentrated solution of potassic hydrate:—

$$\underset{\text{Potassic hydrate.}}{6\mathbf{O}\mathrm{KH}} \quad + \quad \underset{\text{Chlorine.}}{3\mathrm{Cl}_2} \quad = \quad \underset{\text{Potassic chloride.}}{5\mathrm{KCl}} \quad + \quad \underset{\text{Potassic chlorate.}}{\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ko}\end{array}\right.} \quad + \quad \underset{\text{Water.}}{3\mathbf{O}\mathrm{H}_2}.$$

2. Calcic chlorate is made by passing chlorine through boiling milk of lime:—

$$\underset{\text{Calcic Hydrate.}}{6\mathbf{Ca}\mathrm{Ho}_2} \quad + \quad 6\mathrm{Cl}_2 \quad = \quad \underset{\text{Calcic chlorate.}}{\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathrm{Cao}''\\ \mathrm{O}\\ \mathbf{O}\mathrm{Cl}\end{array}\right.} \quad + \quad \underset{\text{Calcic chloride.}}{5\mathbf{Ca}\mathrm{Cl}_2} \quad + \quad \underset{\text{Water.}}{6\mathbf{O}\mathrm{H}_2}.$$

By the addition of potassic chloride to the calcic chlorate, potassic chlorate is formed; the latter is then separated from the calcic chloride by crystallization:—

$$\underset{\text{Calcic chlorate.}}{\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathrm{Cao}''\\ \mathrm{O}\\ \mathbf{O}\mathrm{Cl}\end{array}\right.} + \underset{\text{Potassic chloride.}}{2\mathrm{KCl}} = \underset{\text{Potassic chlorate.}}{2\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ko}\end{array}\right.} + \underset{\text{Calcic chloride.}}{\mathbf{Ca}\mathrm{Cl}_2}.$$

PERCHLORIC ACID.

$$\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ho}\end{array}\right. \text{ or } \left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathrm{O}\\ \mathbf{O}\mathrm{H}\end{array}\right. .$$

Molecular weight = 100·5.

Preparation.—Potassic perchlorate is distilled with about three times its weight of sulphuric acid:—

$$\underset{\text{Potassic perchlorate.}}{2\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ko}\end{array}\right.} + \underset{\text{Sulphuric acid.}}{\mathbf{S}\mathrm{O}_2\mathrm{Ho}_2} = \underset{\text{Perchloric acid.}}{2\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ho}\end{array}\right.} + \underset{\text{Potassic sulphate.}}{\mathbf{S}\mathrm{O}_2\mathrm{Ko}_2}.$$

The impure perchloric acid is then carefully rectified, when pure perchloric acid passes over as an oily liquid towards the end of the operation.

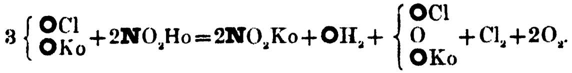
It forms with water a white crystalline hydrate.

Preparation of Potassic Perchlorate.—1. Potassic chlorate is heated gradually, and the process arrested when one third of the oxygen present has been evolved; the residue then contains potassic chloride and perchlorate:—

$$\underset{\text{Potassic chlorate.}}{2\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ko}\end{array}\right.} = \underset{\text{Potassic chloride.}}{\mathrm{KCl}} + \underset{\text{Potassic perchlorate.}}{\left\{\begin{array}{l}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ko}\end{array}\right.} + \mathrm{O}_2.$$

By crystallization, the two salts are separated.

2. When potassic chlorate is gradually introduced into boiling nitric acid, chlorine and oxygen are evolved, potassic nitrate and perchlorate being formed :—



$$3\left\{\begin{matrix}\mathbf{O}\mathrm{Cl}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right. + 2\mathbf{N}\mathrm{O_2Ho} = 2\mathbf{N}\mathrm{O_2Ko} + \mathbf{O}\mathrm{H_2} + \left\{\begin{matrix}\mathbf{O}\mathrm{Cl}\\ \mathrm{O}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right. + \mathrm{Cl_2} + 2\mathrm{O_2}.$$

Potassic chlorate. Nitric acid. Potassic nitrate. Water. Potassic perchlorate.

These salts are then separated by crystallization.

CHAPTER X.

TRIAD ELEMENTS.

SECTION I.

BORON, B_2.

Atomic weight $=11$. *Probable molecular weight* $=22$. *Sp. gr., diamond variety*, 2·68. *Atomicity* ‴. *Evidence of atomicity* :—

Boric chloride	$\mathbf{B}'''\mathrm{Cl}_3$.
Boric fluoride	$\mathbf{B}'''\mathrm{F}_3$.
Boric ethide	$\mathbf{B}'''\mathrm{Et}_3$.

Occurrence.—Found only in combination with oxygen.

Preparation :—

a. Amorphous boron.—1. By igniting boric anhydride with sodium :—

$$\mathbf{B}_2\mathrm{O}_3 \quad + \quad 3\mathrm{Na}_2 \quad = \quad 3\mathbf{O}\mathrm{Na}_2 \quad + \quad \mathrm{B}_2.$$

Boric anhydride. Sodium. Sodic oxide.

2. By passing boric chloride over heated potassium :—

$$\underset{\text{Boric chloride.}}{2\mathbf{B}Cl_3} + 3K_2 = \underset{\text{Potassic chloride.}}{6KCl} + B_2.$$

β. Diamond boron.—By fusing boric anhydride with aluminium :—

$$Al_2 + \underset{\text{Boric anhydride.}}{\mathbf{B}_2O_3} = \underset{\text{Aluminic oxide.}}{\mathbf{Al}'''_2O_3} + B_2.$$

Reactions.—

a. Amorphous boron.—1. Decomposes hot sulphuric acid :—

$$B_2 + \underset{\text{Sulphuric acid.}}{3\mathbf{S}O_2Ho_2} = \underset{\text{Boric anhydride.}}{\mathbf{B}_2O_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} + \underset{\text{Suplhurous anhydride.}}{3\mathbf{S}O_2}.$$

2. Decomposes nitric acid :—

$$B_2 + \underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho} = \underset{\text{Boric acid.}}{2\mathbf{B}Ho_3} + \underset{\text{Nitric peroxide.}}{3\mathbf{N}^{iv}_2O_4}.$$

3. Decomposes alkaline carbonates, sulphates, and nitrates:—

$$B_2 + \underset{\text{Sodic carbonate.}}{3\mathbf{C}ONao_2} = \underset{\text{Trisodic borate.}}{2\mathbf{B}Nao_3} + \underset{\text{Carbonic oxide.}}{3\mathbf{C}''O}.$$

$$B_2 + \underset{\text{Potassic sulphate.}}{3\mathbf{S}O_2Ko_2} = \underset{\text{Tripotassic borate.}}{2\mathbf{B}Ko_3} + \underset{\text{Sulphurous anhydride.}}{3\mathbf{S}O_2}.$$

$$B_2 + \underset{\text{Potassic nitrate.}}{6\mathbf{N}O_2Ko} = \underset{\text{Tripotassic borate.}}{2\mathbf{B}Ko_3} + \underset{\text{Nitric peroxide.}}{3\mathbf{N}^{iv}_2O_4}.$$

4. Boron is one of the few elements which unite directly with nitrogen :—

$$B_2 + N_2 = \underset{\text{Boric nitride.}}{2\mathbf{B}'''N'''}.$$

β. Diamond boron.—1. When fused with hydric potassic sulphate, boric anhydride is formed :—

$$\underset{\text{Hydric potassic sulphate.}}{6\mathbf{S}O_2HoKo} + B_2 = \underset{\text{Boric anhydride.}}{\mathbf{B}_2O_3} + \underset{\text{Potassic sulphate.}}{3\mathbf{S}O_2Ko_2} + \underset{\text{Water.}}{3\mathbf{O}H_2} + \underset{\text{Sulphurous anhydride.}}{3\mathbf{S}O_2}.$$

No compound of boron with hydrogen has been obtained; but the chloride, bromide, and fluoride are known.

BORIC CHLORIDE.

$\mathbf{B}Cl_3$.

Molecular weight =117·5. *Molecular volume* ☐☐. 1 *litre of boric chloride vapour weighs* 58·75 *criths. Sp. gr.* 1·35 *at* 7°. *Boils at* 17°.

Preparation.—By passing chlorine over a mixture of boric anhydride and charcoal heated to redness:—

$$\underset{\text{Boric anhydride.}}{\mathbf{B}_2O_3} + 3Cl_2 + C_3 = \underset{\text{Boric chloride.}}{2\mathbf{B}Cl_3} + \underset{\text{Carbonic oxide.}}{3\mathbf{C}O}.$$

Reaction.—In contact with water it forms hydrochloric and boric acids:—

$$\underset{\text{Boric chloride.}}{\mathbf{B}Cl_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\text{Hydrochloric acid.}}{3HCl} + \underset{\text{Boric acid.}}{\mathbf{B}Ho_3}.$$

BORIC BROMIDE.

$\mathbf{B}Br_3$.

Molecular weight =251. *Molecular volume* ☐☐. 1 *litre of boric bromide vapour weighs* 125·5 *criths. Sp. gr. of liquid* =2·69. *Boils at* 90°.

Prepared and decomposed in exactly the same way as the chloride.

BORIC FLUORIDE.

$\mathbf{B}F_3$.

Molecular weight =68. *Molecular volume* ☐☐. 1 *litre weighs* 34 *criths.*

Preparation.—1. By strongly heating boric anhydride with calcic fluoride:—

$$\underset{\text{Boric anhydride.}}{2\mathbf{B}_2O_3} + \underset{\text{Calcic fluoride.}}{3\mathbf{Ca}F_2} = \underset{\text{Calcic borate.}}{\mathbf{B}_2Cao''_3} + \underset{\text{Boric fluoride.}}{2\mathbf{B}F_3}.$$

2. By heating together boric anhydride with calcic fluoride and sulphuric acid :—

$$\underset{\text{Boric anhydride.}}{\mathbf{B}_2O_3} + \underset{\text{Calcic fluoride.}}{3\mathbf{Ca}F_2} + \underset{\text{Sulphuric acid.}}{3\mathbf{S}O_2Ho_2} = \underset{\text{Dihydric calcic sulphate*.}}{3\mathbf{S}OHo_2Cao''} + \underset{\text{Boric fluoride}}{2\mathbf{B}F_3}.$$

Reaction.—By contact with water, boric fluoride forms a peculiar acid, the hydrofluoboric acid, the constitution of which is not well understood :—

$$\underset{\text{Boric fluoride.}}{4\mathbf{B}F_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\text{Hydrofluoboric acid.}}{3(BF_3,HF)} + \underset{\text{Boric acid}}{\mathbf{B}Ho_3}.$$

This acid acts upon metallic hydrates, forming salts :—

$$\underset{\text{Hydrofluoboric acid.}}{BF_3,HF} + \underset{\text{Potassic hydrate.}}{\mathbf{O}KH} = \underset{\text{Potassic borofluoride}}{BF_3,KF} + \underset{\text{Water}}{\mathbf{O}H_2}.$$

Possibly the boron in these compounds is pentadic—thus, $\mathbf{B}^{v}HF_4$ and $\mathbf{B}^{v}KF_4$.

COMPOUNDS OF BORON WITH OXYGEN AND HYDROXYL.

Boric anhydride	$\mathbf{B}_2O_3$.
Monobasic boric acid Metaboric acid .	$\mathbf{B}OHo$.
Tribasic boric acid . Boric acid	$\mathbf{B}Ho_3$.

BORIC ANHYDRIDE, *Boracic anhydride.*

$\mathbf{B}_2O_3$.

Molecular weight =70. *Sp. gr.* 1·83.

Preparation.—By fusing boric acid at a red heat :—

$$\underset{\text{Boric acid.}}{2\mathbf{B}Ho_3} = \underset{\text{Boric anhydride}}{\mathbf{B}_2O_3} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

---

* See sulphuric acid, Chap. XIII. p. 83.

BORIC ACID, *Boracic Acid, Orthoboric Acid.*

BHo_3.

Molecular weight =62. *Sp. gr.* 1·479.

Occurrence.—Contained in the steam which escapes from the *soffioni* in Tuscany.

Preparation.—By the addition of hydrochloric acid to a hot saturated solution of borax, when the acid crystallizes out on cooling:—

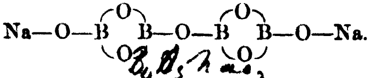
$$\underset{\text{Borax}}{B_4O_5Nao_2} + \underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Water.}}{5OH_2} = \underset{\text{Boric acid.}}{4BHo_3} + \underset{\text{Sodic chloride.}}{2NaCl.}$$

Reactions.—1. At the temperature of 100° it loses water, being converted into metaboric acid:—

$$\underset{\text{Boric acid.}}{BHo_3} = \underset{\text{Metaboric acid.}}{BOHo} + \underset{\text{Water.}}{OH_2}.$$

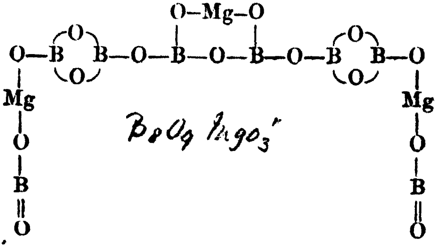
2. By the action of metallic hydrates, oxides, or carbonates, borates are formed.

The mineral *tincal* contains borax, an abnormal sodic borate, $B_4O_5Nao_2$, $10OH_2$.



$B_4O_5Na_2$

A *trimagnesic octoborate* is known as the mineral *boracite*:—



$B_8O_9Mgo_3$

BORIC SULPHIDE. X

$\mathbf{B}_2S''_3$.

Molecular weight =118.

Preparation.—By passing carbonic disulphide over a mixture of carbon and boric anhydride, heated to bright redness:—

$$\underset{\text{Boric anhydride.}}{2\mathbf{B}_2O_3} + \underset{\text{Carbonic disulphide.}}{3\mathbf{C}S''_2} + 3C = \underset{\text{Boric sulphide.}}{2\mathbf{B}_2S''_3} + \underset{\text{Carbonic oxide.}}{6\mathbf{C}''O}.$$

Reaction.—Boric sulphide is readily decomposed by water, giving sulphuretted hydrogen and boric acid:—

$$\underset{\text{Boric sulphide.}}{\mathbf{B}_2S''_3} + \underset{\text{Water.}}{6\mathbf{O}H_2} = \underset{\text{Sulphuretted hydrogen.}}{3\mathbf{S}H_2} + \underset{\text{Boric acid.}}{2\mathbf{B}Ho_3}.$$

BORIC NITRIDE.

$\mathbf{B}N'''$.

Molecular weight =25.

Preparation.—1. By heating boron in nitrogen (see p. 53).
2. By heating together borax and ammonic chloride:—

$$\underset{\text{Borax.}}{\mathbf{B}_4O_5Nao_2} + \underset{\text{Ammonic chloride.}}{4\mathbf{N}H_4Cl} = \underset{\text{Boric nitride.}}{4\mathbf{B}N'''} + \underset{\text{Sodic chloride.}}{2NaCl}$$
$$+ \underset{\text{Water.}}{7\mathbf{O}H_2} + \underset{\text{Hydrochloric acid.}}{2HCl}.$$

Reaction.—When fused with potassic hydrate, boric nitride yields tripotassic borate and ammonia:—

$$\underset{\text{Boric nitride.}}{\mathbf{B}N'''} + \underset{\text{Potassic hydrate.}}{3\mathbf{O}KH} = \underset{\text{Tripotassic borate.}}{\mathbf{B}Ko_3} + \underset{\text{Ammonia.}}{\mathbf{N}H_3}.$$

CHAPTER XI.

TETRAD ELEMENTS.

SECTION I.

CARBON, C.

Atomic weight =12. *Atomicity*$''$ *and*$^{\text{iv}}$. *Evidence of atomicity*:—

Carbonic oxide...........	$\mathbf{C}''O$.	
Carbonic tetrachloride...	$\mathbf{C}^{\text{iv}}Cl_4$.	[illegible]
Marsh-gas.................	$\mathbf{C}^{\text{iv}}H_4$.	[illegible]
Chloroform	$\mathbf{C}^{\text{iv}}HCl$.	[illegible]

Occurrence.—In large quantities in nature, but chiefly in combination.

Three varieties of carbon are known:—

α. *Amorphous carbon.*

Occurrence.—In small quantities in nature, as mineral charcoal.

Preparation.—1. By the action of heat on animal and vegetable matters, without access of air.

2. By the action of potassium, at a high temperature, on carbonic anhydride.

Reaction.—By treatment with a mixture of potassic chlorate and fuming nitric acid, it is converted into brown compounds soluble in water.

β. *Graphite.*

Occurrence.—As the mineral called plumbago.

Preparation.—1. By the action of intense heat on the diamond.

2. By dissolving charcoal in melted cast iron, and subsequent cooling, when the graphite separates in the crystalline state.

3. By heating the various forms of amorphous carbon in the electric arc.

Reaction.—By prolonged digestion with a mixture of nitric acid and potassic chlorate it is converted into *graphitic acid*, a brown crystalline body insoluble in water. Graphitic acid, when heated in the dry state, swells up suddenly, being changed into *pyrographitic oxide,* which is dissolved by the mixture of potassic chlorate and nitric acid.

γ. Diamond.

Occurrence.—In small quantities in nature. Has not yet been obtained artificially.

Reaction.—Is not affected by a mixture of potassic chlorate and nitric acid.

COMPOUNDS OF CARBON WITH OXYGEN.

CARBONIC ANHYDRIDE.

$\mathbf{CO_2}$.

Molecular weight =44. *Molecular volume* ☐☐. 1 *litre weighs* 22 *criths. Fuses at* −57°. *Boils below its melting-point.*

Occurrence.—In the atmosphere, and dissolved in water.

Formation.—By the combustion of carbon and of carbonaceous substances in air or oxygen. In respiration, decay, putrefaction, and fermentation. During the formation of coal. Evolved from volcanoes.

Preparation.—1. By burning carbon in air or oxygen:—

C	+	O_2	=	$\mathbf{CO_2}$.
				Carbonic anhydride.

2. By the action of acids upon metallic carbonates:—

$\mathbf{CO}Ko_2$	+	$\mathbf{SO_2}Ho_2$	=	$\mathbf{CO_2}$	+	$\mathbf{O}H_2$	+	$\mathbf{SO_2}Ko_2$.
Potassic carbonate.		Sulphuric acid.		Carbonic anhydride.		Water.		Potassic sulphate.
$\mathbf{CO}KoHo$	+	$\mathbf{NO_2}Ho$	=	$\mathbf{CO_2}$	+	$\mathbf{O}H_2$	+	$\mathbf{NO_2}Ko$.
Hydric potassic carbonate.		Nitric acid.		Carbonic anhydride.		Water.		Potassic nitrate.
$\mathbf{CO}Cao''$	+	$2HCl$	=	$\mathbf{CO_2}$	+	$\mathbf{O}H_2$	+	$CaCl_2$.
Calcic carbonate.		Hydrochloric acid.		Carbonic anhydride.		Water.		Calcic chloride.

Reactions.—1. Carbonic anhydride is decomposed by heated potassium:—

$$\underset{\text{Carbonic anhydride.}}{3CO_2} + 2K_2 = \underset{\text{Potassic carbonate.}}{2COKo_2} + C.$$

2. It acts upon metallic hydrates, forming carbonates:—

$$\underset{\text{Carbonic anhydride.}}{CO_2} + \underset{\text{Potassic hydrate.}}{2KHo} = \underset{\text{Potassic carbonate.}}{COKo_2} + \underset{\text{Water.}}{OH_2}.$$

$$\underset{\text{Carbonic anhydride.}}{CO_2} + \underset{\text{Calcic hydrate.}}{CaHo_2} = \underset{\text{Calcic carbonate.}}{COCao''} + \underset{\text{Water.}}{OH_2}.$$

Carbonic acid, $COHo_2$, is not known.

CARBONIC OXIDE.

CO.

Molecular weight =28. *Molecular volume* ◫. 1 *litre weighs* 14 *criths.*

Formation.—In the combustion of carbon or carbonaceous matter, with a limited supply of air. In the destructive distillation of many organic substances containing oxygen.

Preparation.—1. By passing carbonic anhydride over red-hot charcoal:—

$$\underset{\text{Carbonic anhydride.}}{CO_2} + C = \underset{\text{Carbonic oxide.}}{2CO}.$$

2. By passing carbonic anhydride over red-hot iron:—

$$\underset{\text{Carbonic anhydride.}}{4CO_2} + Fe_3 = \underset{\text{Triferric tetroxide.}}{{}^{iv}(Fe_3)^{viii}O_4} + \underset{\text{Carbonic oxide.}}{4CO}.$$

3. By heating iron or carbon with a carbonate:—

$$\underset{\text{Calcic carbonate.}}{COCao''} + C = \underset{\text{Lime.}}{CaO} + \underset{\text{Carbonic oxide.}}{2CO}.$$

4. By heating oxalic acid with sulphuric acid (by which water is removed from the former), and then separating the carbonic anhydride by washing with sodic hydrate:—

$$\underset{\text{Oxalic acid.}}{\left\{\begin{matrix}COHo\\COHo\end{matrix}\right.} = \underset{\text{Water.}}{OH_2} + \underset{\text{Carbonic oxide.}}{CO} + \underset{\text{Carbonic anhydride.}}{CO_2}.$$

5. By heating formic acid, or a formate, with sulphuric acid:—

$$\left\{\begin{matrix}\mathrm{H} \\ \mathbf{C}\mathrm{OHo}\end{matrix}\right. = \mathbf{O}\mathrm{H}_2 + \mathbf{C}\mathrm{O}.$$

Formic acid. Water. Carbonic oxide.

6. By heating potassic ferrocyanide with sulphuric acid:—

$$\mathrm{Fe}''\mathrm{C}_6\mathrm{N}_6\mathrm{K}_4 + 6\mathbf{O}\mathrm{H}_2 + 6\mathbf{S}\mathrm{O}_2\mathrm{Ho}_2 = 6\mathbf{C}\mathrm{O}$$

Potassic ferrocyanide. Water. Sulphuric acid. Carbonic oxide.

$$+ 2\mathbf{S}\mathrm{O}_2\mathrm{Ko}_2 + \mathbf{S}\mathrm{O}_2\mathrm{Feo}'' + 3\mathbf{S}\mathrm{O}_2(\mathrm{NH}_4\mathrm{O})_2.$$

Potassic sulphate. Ferrous sulphate. Ammonic sulphate.

Reactions.—1. It burns in air and oxygen, producing carbonic anhydride:—

$$\mathbf{C}\mathrm{O} + \mathrm{O} = \mathbf{C}\mathrm{O}_2.$$

Carbonic oxide. Carbonic anhydride.

2, Carbonic oxide and chlorine unite under the influence of light (p. 30), forming *carbonic oxydichloride*, or *phosgene gas*, $\mathbf{C}\mathrm{OCl}_2$.

The compounds of carbon with chlorine, nitrogen, and hydrogen will be studied in connexion with organic compounds.

CHAPTER XII.

PENTAD ELEMENTS.

SECTION I.

NITROGEN, *Azote*, N_2.

***Atomic weight* = 14. *Molecular weight* = 28. *Molecular volume* ☐☐. 1 *litre weighs* 14 *criths. Atomicity* $^{\mathrm{v}}$, *which, by the mutual saturation of pairs of bonds, becomes reduced to* $'''$ *or to* $'$ (see p. 20). *Evidence of atomicity*:—**

Nitrous oxide..........	$\mathbf{O}\mathrm{N}_2$.
Ammonia	$\mathbf{N}'''\mathrm{H}_3$.
Ammonic chloride	$\mathbf{N}^{\mathrm{v}}\mathrm{H}_4\mathrm{Cl}$.

Occurrence.—In the free state in the atmosphere. In some nebulæ? In combination, in animal and vegetable bodies, and in native nitrates.

Preparation.—1. By burning phosphorus in air, whereby the oxygen is removed from the latter.

2. By passing air over ignited copper, when the oxygen unites with the copper.

3. By heating ammonic nitrite, or a mixture of ammonic chloride and potassic or sodic nitrite:—

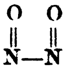
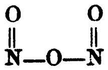
$$\underset{\text{Ammonic nitrite.}}{\mathbf{N}'''O(N^{v}H_4O)} = N_2 + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

$$\underset{\text{Ammonic chloride.}}{\mathbf{N}H_4Cl} + \underset{\text{Sodic nitrite.}}{\mathbf{N}ONao} = \underset{\text{Sodic chloride.}}{NaCl} + N_2 + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

4. By passing chlorine through an excess of solution of ammonia:—

$$\underset{\text{Ammonia.}}{8\mathbf{N}H_3} + 3Cl_2 = \underset{\text{Ammonic chloride.}}{6\mathbf{N}H_4Cl} + N.$$

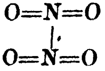
COMPOUNDS OF NITROGEN WITH OXYGEN AND HYDROXYL.

Nitrous oxide..	$\mathbf{O}N_2$.	N—O—N
Nitric oxide*	$\left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right.$	 $\begin{matrix}O & & O\\ \Vert & & \Vert\\ N & — & N\end{matrix}$
Nitrous anhydride	$\left\{\begin{matrix}\mathbf{N}O\\ O\\ \mathbf{N}O\end{matrix}\right.$.	 $\begin{matrix}O & & & & O\\ \Vert & & & & \Vert\\ N & — & O & — & N\end{matrix}$

---

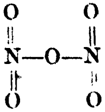
* This compound is anomalous; for its molecule, deduced from the specific gravity, is represented by NO. The dissociation, which in the case of $\left\{\begin{matrix}NO_2\\ NO_2\end{matrix}\right.$ is very imperfect at 0° C., but almost complete at 100° C., is probably nearly complete in the case of N_2O_2 at the lowest temperature to which this gas has hitherto been exposed.

Nitric peroxide .... $\left\{\begin{array}{l}\mathbf{N}O_2\\ \mathbf{N}O_2\end{array}\right.$



$\begin{array}{c}O{=}N{=}O\\ |\\ O{=}N{=}O\end{array}$

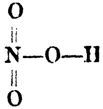
Nitric anhydride ...... $\left\{\begin{array}{l}\mathbf{N}O_2\\ O\\ \mathbf{N}O_2\end{array}\right.$



$\begin{array}{ccccc}O&&&&O\\ \|&&&&\|\\ N&-&O&-&N\\ \|&&&&\|\\ O&&&&O\end{array}$

Nitrous acid .... $\mathbf{N}OHo$. $O{=}N{-}O{-}H$

Nitric acid .... $\mathbf{N}O_2Ho$.



$\begin{array}{lll}O&&\\ \|&&\\ N&-O-&H\\ \|&&\\ O&&\end{array}$

NITRIC ACID, *Aquafortis*.

$\mathbf{N}O_2Ho$.

Molecular weight = 63. *Molecular volume* ☐☐. 1 *litre of nitric acid vapour weighs* 31·5 *criths. Fuses at* −50°. *Boils at* 84°·5.

Production.—1. By the slow oxidation of nitrogenized organic matter, in the presence of powerful bases.

2. By the passage of electric sparks through moist air.

Manufacture.—By distilling potassic nitrate (nitre), or sodic nitrate (cubic nitre), with concentrated sulphuric acid:—

$$\underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Hydric potassic sulphate.}}{\mathbf{S}O_2HoKo} + \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho}.$$

By employing two molecules of potassic nitrate and one of sulphuric acid a saving of sulphuric acid is effected, but a higher temperature is required, which destroys some of the nitric acid. The reaction takes place in two stages:—

$$\text{I. } \underset{\text{Potassic nitrate.}}{2\mathbf{N}O_2Ko} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Hydric potassic sulphate.}}{\mathbf{S}O_2HoKo}$$

$$+ \underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko} + \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho}.$$

$$2.\ \underset{\text{Hydric potassic sulphate.}}{\mathbf{S}O_2HoKo} + \underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko} = \underset{\text{Potassic sulphate.}}{\mathbf{S}O_2Ko_2} + \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho.}$$

Decompositions.—1. The decomposition which nitric acid undergoes by heat is expressed in the following equation:—

$$\underset{\text{Nitric acid.}}{4\mathbf{N}O_2Ho} = \underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Nitric peroxide.}}{2'\mathbf{N}^{iv}{}_2O_4} + \underset{\text{Oxygen.}}{O_2}$$

2. By the action of metallic oxides or hydrates, nitric acid produces nitrates:—

$$\underset{\text{Potassic hydrate.}}{\mathbf{O}KH} + \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho} = \underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko} + \underset{\text{Water.}}{\mathbf{O}H_2.}$$

$$\underset{\text{Plumbic oxide.}}{\mathbf{Pb}O} + \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} = \underset{\text{Plumbic nitrate.}}{\left\{\begin{matrix}\mathbf{N}O_2\\ Pbo''\\ \mathbf{N}O_2\end{matrix}\right.} + \underset{\text{Water.}}{\mathbf{O}H_2.}$$

NITRIC ANHYDRIDE.

$\mathbf{N}_2O_5$.

Probable molecular weight =108. *Probable molecular volume* ☐☐. *Fuses at* 29°·5. *Boils at* 45°.

Preparation.—1. By passing dry chlorine over argentic nitrate, heated first to 93° and then to 66°.

$$\underset{\text{Argentic nitrate.}}{4\mathbf{N}O_2Ago} + 2Cl_2 = \underset{\text{Argentic chloride.}}{4AgCl} + \underset{\text{Nitric anhydride.}}{2\mathbf{N}_2O_5} + O_2.$$

2. By passing the vapour of nitric dioxychloride over argentic nitrate heated to 60°–70°.

$$\underset{\text{Nitric dioxychloride.}}{\mathbf{N}O_2Cl} + \underset{\text{Argentic nitrate.}}{\mathbf{N}O_2Ago} = \underset{\text{Argentic chloride.}}{AgCl} + \underset{\text{Nitric anhydride.}}{\mathbf{N}_2O_5.}$$

Reaction.—By the action of water it forms nitric acid:—

$$\underset{\text{Nitric anhydride.}}{\mathbf{N}_2O_5} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho.}$$

NITROUS ANHYDRIDE.

$\mathbf{N}_2O_3$.

Probable molecular weight =76. *Probable molecular volume* $\square\square$.

Preparation.—1. By heating together nitric acid and starch.

2. By gently heating nitric acid with arsenious anhydride:—

$$\underset{\text{Arsenious anhydride.}}{\mathbf{As}_2O_3} + \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} = \underset{\text{Arsenic anhydride.}}{\mathbf{As}_2O_5} + \underset{\text{Nitrous anhydride.}}{\mathbf{N}_2O_3} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

3. By the action of strong nitric acid on silver:—

$$\underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho} + 2Ag_2 = \underset{\text{Argentic nitrate.}}{4\mathbf{N}O_2Ago} + \underset{\text{Nitrous anhydride.}}{\mathbf{N}_2O_3} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

NITROUS ACID.

$\mathbf{N}OHo$.

Molecular weight =47.

Preparation.—By mixing liquefied nitrous anhydride with a small quantity of water:—

$$\underset{\text{Nitrous anhydride.}}{\mathbf{N}_2O_3} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Nitrous acid.}}{2\mathbf{N}OHo}.$$

Decompositions.—1. In the presence of much water, nitric acid and nitric oxide are formed:—

$$\underset{\text{Nitrous acid.}}{3\mathbf{N}OHo} = \underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho} + \underset{\text{Nitric oxide.}}{\mathbf{N}''_2O_2} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

2. Nitrous acid acts as a reducing agent under some circumstances:—

$$\underset{\text{Nitrous acid.}}{2\mathbf{N}OHo} + O_2 = \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho};$$

and as an oxidizing agent under others:—

$$\underset{\text{Nitrous acid.}}{4\mathbf{N}OHo} = \underset{\text{Nitric oxide.}}{2\mathbf{N}''_2O_2} + \underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Oxygen.}}{O_2}.$$

3. By the action of metallic oxides or hydrates, nitrous acid forms nitrites:—

$$\underset{\text{Potassic hydrate.}}{\mathbf{O}KH} + \underset{\text{Nitrous acid.}}{\mathbf{N}OHo} = \underset{\text{Potassic nitrite.}}{\mathbf{N}OKo} + \underset{\text{Water.}}{\mathbf{O}H}.$$

NITROUS OXIDE, *Laughing Gas.*

$$ON_2.$$

Molecular weight $=44$. ***Molecular volume*** □□. 1 ***litre weighs*** 22 ***criths.*** ***Fuses at*** $-101°$. ***Boils at*** $-88°$.

Preparation.—1. By the action of dilute nitric acid on zinc:—

$$\underset{\text{Nitric acid.}}{10\mathbf{N}O_2Ho} + Zn_4 = \underset{\text{Nitrous oxide.}}{ON_2} + \underset{\text{Zincic Nitrate.}}{4\left\{\begin{matrix}\mathbf{N}O_2\\ Zno''\\ \mathbf{N}O_2\end{matrix}\right.} + \underset{\text{Water.}}{5OH_2}.$$

2. By heating ammonic nitrate:—

$$\underset{\text{Ammonic nitrate.}}{\mathbf{N}O_2(N^{v}H_4O)} = \underset{\text{Water.}}{2OH_2} + \underset{\text{Nitrous oxide.}}{ON_2}.$$

NITRIC OXIDE.

$$\left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right., \text{ or } {}'\mathbf{N}''_2O_2.$$

Molecular weight $=60$. *Molecular volume anomalous* 1 *litre weighs* 15 *criths.*

Preparation.—By the action of nitric acid upon mercury or copper:—

$$3Cu + \underset{\text{Nitric acid.}}{8\mathbf{N}O_2Ho} = \underset{\text{Cupric nitrate.}}{3\left\{\begin{matrix}\mathbf{N}O_2\\ Cuo''\\ \mathbf{N}O_2\end{matrix}\right.} + \underset{\text{Nitric oxide.}}{{}'\mathbf{N}''_2O_2} + \underset{\text{Water.}}{4OH_2}.$$

Reaction.—Unites directly with oxygen:—

$$\underset{\text{Nitric oxide.}}{2{}'\mathbf{N}''_2O_2} + O_2 = \underset{\text{Nitrous anhydride.}}{2\mathbf{N}'''_2O_3}:$$

$$\underset{\text{Nitric oxide.}}{{}'\mathbf{N}''_2O_2} + O_2 = \underset{\text{Nitric peroxide.}}{{}'\mathbf{N}^{iv}_2O_4}.$$

NITRIC PEROXIDE.

$$\left\{\begin{matrix}\mathbf{N}O_2\\ \mathbf{N}O_2\end{matrix}\right., \text{ or } \mathbf{N}^{iv}{}_2O_4.$$

Molecular weight = 92. *Molecular volume* ◫ *to* ⊞. 1 *litre weighs* 23 *to* 46 *criths. Melts at* −9°. *Boils at* 22°.

Preparation.—1. By the union of nitric oxide with oxygen (see above).

2. By the action of nitric acid upon tin :—

$$Sn_5 + \underset{\text{Nitric acid.}}{20\mathbf{N}O_2Ho} = \underset{\text{Metastannic acid.}}{\mathbf{Sn}_5O_5Ho_{10}} + \underset{\text{Water.}}{5\mathbf{O}H_2} + \underset{\text{Nitric peroxide.}}{10\mathbf{N}^{iv}{}_2O_4}.$$

Decomposition.—By the action of metallic hydrates and oxides it produces nitrites and nitrates :—

$$\underset{\text{Nitric peroxide.}}{\mathbf{N}^{iv}{}_2O_4} + \underset{\text{Potassic hydrate.}}{2\mathbf{O}KH} = \underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko} + \underset{\text{Potassic nitrite.}}{\mathbf{N}OKo} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

COMPOUNDS CONTAINING NITROGEN, CHLORINE, AND OXYGEN.

NITROUS OXYCHLORIDE, *Chloronitrous Gas.*

$$\mathbf{N}OCl.$$

Molecular weight = 65·5. *Molecular volume* ◫. 1 *litre weighs* 32·75 *criths. Boils at* 0°.

A mixture of nitric and hydrochloric acids possesses the property of dissolving gold, and is therefore called aqua regia; when heated it evolves chlorine and nitrous oxychloride :—

$$\underset{\text{Nitric acid.}}{\mathbf{N}O_2Ho} + \underset{\text{Hydrochloric acid.}}{3HCl} = \underset{\text{Nitrous oxychloride.}}{\mathbf{N}OCl} + \underset{\text{Water.}}{2\mathbf{O}H_2} + Cl_2.$$

NITRIC DIOXY-TETRACHLORIDE, *Chloronitric Gas.*

$$'\mathbf{N}^{iv}{}_2O_2Cl_4.$$

Prepared, together with nitrous oxychloride, by heating a mixture of nitric and hydrochloric acids :—

$$\underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Chloronitric gas.}}{'\mathbf{N}^{iv}{}_2O_2Cl_4} + \underset{\text{Water.}}{4\mathbf{O}H_2} + Cl_2.$$

NITRIC DIOXYCHLORIDE, *Chloropernitric Gas.*

$$\mathbf{N}O_2Cl.$$

Preparation.—By mixing phosphoric oxytrichloride and plumbic nitrate :—

$$\underset{\text{Plumbic nitrate.}}{3\left\{\begin{matrix}\mathbf{N}O_2\\ Pbo''\\ \mathbf{N}O_2\end{matrix}\right.} + \underset{\text{Phosphoric oxytrichloride.}}{2\mathbf{P}Cl_3O} = \underset{\text{Triplumbic diphosphate.}}{\mathbf{P}_2O_2Pbo''_3} + \underset{\text{Nitric dioxychloride.}}{6\mathbf{N}O_2Cl.}$$

COMPOUNDS OF NITROGEN WITH HYDROGEN.

AMMONIA.

$$\mathbf{N}H_3.$$

Molecular weight = 17. *Molecular volume* ☐☐. 1 *litre weighs* 8·5 *criths. Fuses at* −75°. *Boils at* −38°·5.

Occurrence.—In the atmosphere and in rain water in very minute quantities.

Formation.—By the decay of animal and vegetable matters containing nitrogen.

Manufacture—By the destructive distillation of animal matter, as horn or bones, and of vegetable matter, as coal.

Preparation.—By heating a mixture of lime and ammonic chloride (sal-ammoniac) :—

$$\underset{\text{Ammonic chloride.}}{2\mathbf{N}H_4Cl} + \underset{\text{Lime.}}{\mathbf{Ca}O} = \underset{\text{Calcic chloride.}}{\mathbf{Ca}Cl_2} + \underset{\text{Ammonia.}}{2\mathbf{N}H_3} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

Reactions.—1. Decomposed by chlorine (see p. 62).

2. Unites directly with acids, forming the ammonium salts in which the atomicity of nitrogen is ^{v}:—

$$\mathbf{N}'''H_3 + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Ammonic chloride *.}}{\mathbf{N}^vH_4Cl.}$$

$$\mathbf{N}'''H_3 + \underset{\text{Nitric acid.}}{\mathbf{N}^vO_2Ho} = \underset{\text{Ammonic nitrate †.}}{\mathbf{N}^vO_2(N^vH_4O).}$$

$$2\mathbf{N}'''H_3 + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Ammonic sulphate ‡.}}{\mathbf{S}O_2(N^vH_4O)_2.}$$

AMMONIUM.

$$\left\{\begin{matrix}\mathbf{N}H_4\\ \mathbf{N}H_4'\end{matrix}\right.$$

This monad radical has never been obtained in the free state; but its compounds are perfectly analogous, in crystalline form and other properties, to those of potassium. These facts have induced some chemists to consider the group NH_4 as a metal, to which they have given the name ammonium—an hypothesis which is considered to receive support from the production of an unstable amalgam of this radical. All the compounds of mercury with metals are found to possess metallic lustre; and this is also the case with the amalgam of ammonium. It may be prepared by two different processes.

1. If a solution of ammonic chloride be electrolyzed, the negative electrode being mercury and the positive a platinum

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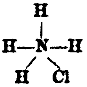

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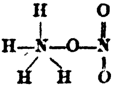

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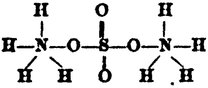


plate, the mercury is observed to swell up, owing to the formation of a spongy metallic mass.

2. By preparing an amalgam of potassium or of sodium, and pouring it into a slightly warmed solution of ammonic chloride, the amalgam is found to swell enormously, potassic or sodic chloride being simultaneously formed:—

$$\underset{\text{Sodic amalgam.}}{Hg_nNa_m} + \underset{\text{Ammonic chloride.}}{m\mathbf{N}H_4Cl} = \underset{\text{Ammonic amalgam.}}{Hg_n(N^vH_4)_m} + \underset{\text{Sodic chloride.}}{mNaCl.}$$

Ammonic amalgam rapidly decomposes into mercury, ammonia, and hydrogen, the ammonia and hydrogen being liberated in the proportions of $2NH_3$ to H_2:—

$$\underset{\text{Ammonic amalgam.}}{2Hg_n(N^vH_4)_m} = \underset{\text{Mercury.}}{2nHg} + \underset{\text{Ammonia.}}{2m\mathbf{N}H_3} + mH_2.$$

Ammonium plays the part of a compound monad radical; and its salts are isomorphous with those of potassium; they are all volatile, unless the acid from which they are derived be fixed.

COMPOUND OF NITROGEN WITH CHLORINE.

NITROUS CHLORIDE.

$\mathbf{N}Cl_3$?

Preparation.—By the action of chlorine upon ammonic chloride:—

$$\underset{\text{Ammonic chloride.}}{\mathbf{N}^vH_4Cl} + 3Cl_2 = \underset{\text{Nitrous chloride.}}{\mathbf{N}''Cl_3} + \underset{\text{Hydrochloric acid.}}{4HCl.}$$

The formula of this compound is not fixed with certainty: it may contain hydrogen; and it is possible that the two compounds intermediate between ammonia and nitrous chloride may exist:—

$$\mathbf{N}H_3,\ \mathbf{N}H_2Cl,\ \mathbf{N}HCl_2,\ \mathbf{N}Cl_3.$$

COMPOUND OF NITROGEN WITH IODINE AND HYDROGEN.

NITROUS HYDRODINIODIDE.

Preparation.—By the action of ammonia on iodine a brown substance is obtained, which has the composition $\mathbf{N}HI_2$. It is formed according to the following equation:—

$$\underset{\text{Ammonia.}}{3\mathbf{N}H_3} + 2I_2 = \underset{\text{Nitrous hydrodiniodide.}}{\mathbf{N}HI_2} + \underset{\text{Ammonic iodide.}}{2\mathbf{N}H_4I}.$$

CHAPTER XIII.

HEXAD ELEMENTS.

SECTION I.

SULPHUR, S_2.

Atomic weight = 32. *Molecular weight* = 64. *Molecular volume* ☐☐ *at* 1000°, *but only one-third of this at its boiling-point.* 1 *litre of sulphur vapour weighs* 32 *criths. Rhomboidal variety fuses at* 114°·5 *and boils at* 445°. *Atomicity* $''$ ^{iv} *and* ^{vi}. *Evidence of atomicity*:—

Hydrosulphuric acid	$\mathbf{S}''H_2$.
Triethylsulphine iodide	$\mathbf{S}^{iv}Et_3I$.
Sulphuric dioxydichloride	$\mathbf{S}^{vi}O_2Cl_2$.
Sodic nitrosulphate	$\mathbf{S}^{vi}O(NO)_2Nao_2$.

Occurrence.—Found in the free state in volcanic districts, and widely diffused in combination with metals and oxygen, as sulphides and sulphates.

Manufactured from native sulphur, and from

Iron pyrites	$\mathbf{Fe}S''_2$.
Copper pyrites	$(\mathbf{FeCu})S''_2$.
Calcic oxysulphide	$\mathbf{Ca}_2OS''$.

Character.—Sulphur is capable of existing in several allotropic forms, of which the following are the most important:—

	Condition.	Specific gravity.	Behaviour with carbonic disulphide.
α.	Octahedral	2·05	Soluble.
β.	Prismatic	1·98	Transformed into α.
γ.	Plastic	1·95	Insoluble.
δ.	Powder	1·95	Insoluble.

When united exclusively with positive elements or radicals, sulphur is almost invariably a dyad; and it is then the analogue of oxygen, as will be seen from the following formulæ:—

Oxygen compounds ... $\mathbf{O}K_2$, $\mathbf{O}KH$, $\mathbf{C}O_2$, $\mathbf{C}OKo_2$.
Sulphur compounds ... $\mathbf{S}K_2$, $\mathbf{S}KH$, $\mathbf{C}S''_2$, $\mathbf{C}S''Ks_2$.

COMPOUNDS OF SULPHUR WITH POSITIVE ELEMENTS.

Sulphuretted hydrogen	$\mathbf{S}H_2$.
Hydrosulphyl	$'\mathbf{S}'_2H_2$, or Hs_2.
Carbonic disulphide	$\mathbf{C}S''_2$.

SULPHURETTED HYDROGEN, *Hydrosulphuric Acid, Sulphhydric Acid.*

$\mathbf{S}H_2$. H—S—H.

***Molecular weight* =34. *Molecular volume* ☐☐. 1 *litre weighs* 17 *criths.* *Solid at* $-85^\circ\cdot5$. *Liquid under a pressure of* 17 *atmospheres at* 10°.**

Occurrence.—Evolved with other gases from volcanoes and fumaroles. Found also in hepatic mineral waters, and frequently in waters which contain both organic matters and sulphates.

Preparation.—1. By direct union of its elements:—

$$H_2 + S = \mathbf{S}H_2.$$

2. By the action of hydrochloric or dilute sulphuric acid on ferrous sulphide:—

$$\underset{\substack{\text{Ferrous}\\\text{sulphide.}}}{\mathbf{Fe}S''} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{2HCl} = \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{\mathbf{S}H_2} + \underset{\substack{\text{Ferrous}\\\text{chloride.}}}{\mathbf{Fe}Cl_2}.$$

$$\underset{\substack{\text{Ferrous}\\\text{sulphide.}}}{\mathbf{Fe}S''} + \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{\mathbf{S}O_2Ho_2} = \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{\mathbf{S}H_2} + \underset{\substack{\text{Ferrous}\\\text{sulphate.}}}{\mathbf{S}O_2Feo''}.$$

3. By the action of hydrochloric acid on antimonious sulphide with the aid of a gentle heat:—

$$\underset{\substack{\text{Antimonious}\\\text{sulphide.}}}{\mathbf{Sb}_2S''_3} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{6HCl} = \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{3\mathbf{S}H_2} + \underset{\substack{\text{Antimonious}\\\text{chloride.}}}{2\mathbf{Sb}Cl_3}.$$

Reactions.—1. It is immediately decomposed by chlorine, thus:—

$$\mathbf{S}H_2 + Cl_2 = 2HCl + S.$$

2. It is also rapidly decomposed by many metallic compounds rich in oxygen, such as ferric oxide:—

$$\underset{\text{Ferric oxide.}}{\mathbf{Fe}'''_2O_3} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{3\mathbf{S}H_2} = \underset{\substack{\text{Ferrous}\\\text{sulphide.}}}{2\mathbf{Fe}S''} + S + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

3. The sulphhydrates and sulphides of the metals are produced by the action of hydrosulphuric acid on the hydrates and oxides, thus:—

$$\underset{\substack{\text{Potassic}\\\text{hydrate.}}}{\mathbf{O}KH} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{\mathbf{S}H_2} = \underset{\substack{\text{Potassic}\\\text{sulphhydrate.}}}{\mathbf{S}KH} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

$$\underset{\text{Baric hydrate.}}{\mathbf{Ba}Ho_2} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{2\mathbf{S}H_2} = \underset{\substack{\text{Baric}\\\text{sulphhydrate.}}}{\mathbf{Ba}Hs_2} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

$$\underset{\substack{\text{Argentic}\\\text{oxide.}}}{\mathbf{O}Ag_2} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{\mathbf{S}H_2} = \underset{\substack{\text{Argentic}\\\text{sulphide.}}}{\mathbf{S}Ag_2} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

$$\underset{\substack{\text{Cupric}\\\text{oxide.}}}{\mathbf{Cu}O} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{\mathbf{S}H_2} = \underset{\substack{\text{Cupric}\\\text{sulphide.}}}{\mathbf{Cu}S''} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

HYDROSULPHYL, *Hydric Persulphide.*

$$'\mathbf{S}'_2H_2, \text{ or } Hs_2.$$

$$H—S—S—H.$$

Probable molecular weight =66. *Sp. gr.* 1·769.

Preparation.—By pouring a solution of calcic disulphide into hydrochloric acid:—

$$\underset{\text{Calcic disulphide.}}{'\mathbf{S}'_2Ca''} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Hydrosulphyl.}}{'\mathbf{S}'_2H_2} + \underset{\text{Calcic chloride.}}{\mathbf{Ca}Cl_2}.$$

Character.—It is the analogue of hydroxyl in composition and functions.

HYPO-SULPHUROUS HYDROSULPHATE.

$$\mathbf{S}Hs_2 \text{ or } \begin{cases} Hs \\ S'' \\ Hs \end{cases}$$

Probable molecular weight = 98.

Preparation.—When a cold saturated solution of strychnine in alcohol is mixed with an alcoholic solution of yellow ammonic sulphide, a crystalline compound is formed containing $C_{21}H_{22}N_2O_2, H_2S_2$. By the action of sulphuric acid upon this compound, hyposulphurous hydrosulphate is liberated as a yellow oily body.

CARBONIC DISULPHIDE, *Bisulphide of Carbon.*

$$\mathbf{C}S_2.$$

Molecular weight =76. *Molecular volume* ☐☐. 1 *litre of carbonic disulphide vapour weighs* 38 *criths. Specific gravity of liquid* 1·293. *Boils at* 46°·6.

Preparation.—1. By passing sulphur over strongly ignited charcoal:—

$$C + S_2 = \underset{\text{Carbonic disulphide.}}{\mathbf{C}S''_2}.$$

2. By heating together charcoal and iron- or copper-pyrites :—

$$\mathbf{C} + \underset{\substack{\text{Iron-pyrites.}\\ \text{Ferric disulphide.}}}{2\mathbf{Fe}S''_2} = \underset{\substack{\text{Carbonic}\\ \text{disulphide.}}}{\mathbf{C}S''_2} + \underset{\substack{\text{Ferrous}\\ \text{sulphide.}}}{2\mathbf{Fe}S''}.$$

Decompositions.—1. Heated potassium burns in the vapour of carbonic disulphide, with formation of potassic sulphide and liberation of carbon :—

$$\underset{\text{Carbonic disulphide.}}{\mathbf{C}S''_2} + 2K_2 = \underset{\text{Potassic sulphide.}}{2\mathbf{S}K_2} + C.$$

2. When brought into contact with a solution of an alkaline hydrate, carbonic disulphide is decomposed, a carbonate and a sulpho-carbonate being formed :—

$$\underset{\substack{\text{Potassic}\\ \text{hydrate.}}}{6\mathbf{O}KH} + \underset{\substack{\text{Carbonic}\\ \text{disulphide.}}}{3\mathbf{C}S''_2} = \underset{\substack{\text{Potassic}\\ \text{sulpho-carbonate.}}}{2\mathbf{C}S''Ks_2} + \underset{\substack{\text{Potassic}\\ \text{carbonate.}}}{\text{\c{C}}OKo_2} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

3. In contact with solutions of alkaline sulphides, carbonic disulphide also forms alkaline sulpho-carbonates :—

$$\underset{\text{Potassic sulphide.}}{\mathbf{S}K_2} + \underset{\text{Carbonic disulphide.}}{\mathbf{C}S''_2} = \underset{\text{Potassic sulpho-carbonate.}}{\mathbf{C}S''Ks_2}.$$

4. When the vapour of carbonic disulphide is passed over heated calcic hydrate it is decomposed, carbonic anhydride and sulphuretted hydrogen being evolved :—

$$\underset{\substack{\text{Carbonic}\\ \text{disulphide.}}}{\mathbf{C}S_2} + \underset{\substack{\text{Calcic}\\ \text{hydrate.}}}{2\mathbf{Ca}Ho_2} = \underset{\substack{\text{Calcic}\\ \text{oxide.}}}{2\mathbf{Ca}O} + \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{\mathbf{C}O_2} + \underset{\substack{\text{Sulphuretted}\\ \text{hydrogen.}}}{2\mathbf{S}H_2}.$$

SULPHO-CARBONIC ACID.

$$\mathbf{C}S''Hs_2.$$

Preparation.—By the action of hydrochloric acid on ammonic sulpho-carbonate :—

$$\underset{\substack{\text{Ammonic}\\ \text{sulpho-carbonate.}}}{\mathbf{C}S''(NH_4S)_2} + \underset{\substack{\text{Hydrochloric}\\ \text{acid.}}}{2HCl} = \underset{\substack{\text{Sulpho-carbonic}\\ \text{acid.}}}{\mathbf{C}S''Hs_2} + \underset{\substack{\text{Ammonic}\\ \text{chloride.}}}{2\mathbf{N}H_4Cl}.$$

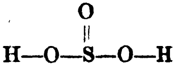
COMPOUNDS OF SULPHUR WITH OXYGEN AND HYDROXYL.

In these compounds the sulphur is either a dyad, a tetrad, or a hexad.

Hyposulphurous acid ...... $\mathbf{S}Ho_2$. H—O—S—O—H

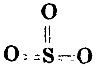
Sulphurous anhydride ...... $\mathbf{S}O_2$. O=S=O

Sulphurous acid .... $\mathbf{S}OHo_2$.



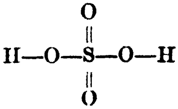
$$H—O—\overset{\overset{O}{\|}}{S}—O—H$$

Sulphuric anhydride........ $\mathbf{S}O_3$.



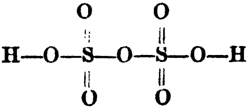
$$O=\overset{\overset{O}{\|}}{S}=O$$

Sulphuric acid. (*Hydric sulphate.*) .... $\mathbf{S}O_2Ho_2$.



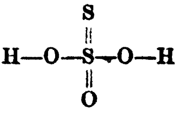
$$H—O—\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}—O—H$$

Nordhausen sulphuric acid. (*Dihydric disulphate.*) .... $\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ O\\ \mathbf{S}O_2Ho\end{array}\right.$.



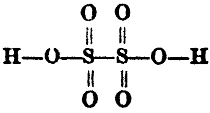
$$H—O—\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}—O—\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}—O—H$$

Thiosulphuric acid .... $\mathbf{S}S''OHo_2$.



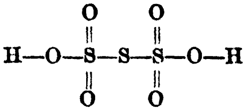
$$H—O—\underset{\underset{O}{\|}}{\overset{\overset{S}{\|}}{S}}—O—H$$

Dithionic acid ... $'\mathbf{S}^{v}{}_2O_4Ho_2$, or $\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ \mathbf{S}O_2Ho\end{array}\right.$.



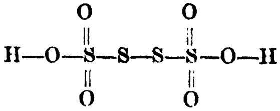
$$H—O—\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}—\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}—O—H$$

Trithionic acid. *Sulphodithionic acid*) ... $\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ \mathbf{S}O_2Ho\end{array}\right.$.



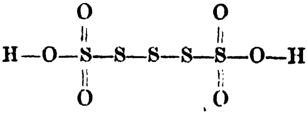
$$H—O—\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}—S—\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}—O—H$$

Tetrathionic acid. (*Disulphodithionic acid.*) ...... $\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ S''\\ \mathbf{S}O_2Ho\end{array}\right.$.



$$H—O—\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}—S—S—\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}—O—H$$

Pentathionic acid. (*Trisulphodithionic acid.*) ...... $\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ S''\\ S''\\ \mathbf{S}O_2Ho\end{array}\right.$.



$$H—O—\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}—S—S—S—\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}—O—H$$

HYPOSULPHUROUS ACID.

$\mathbf{S}Ho_2$.

Formation.—By the action of zinc upon aqueous sulphurous acid :—

$$\underset{\text{Sulphurous acid.}}{\mathbf{S}OHo_2} + Zn = \underset{\text{Hyposulphurous acid.}}{\mathbf{S}Ho_2} + ZnO.$$

Reaction.—Absorbs free oxygen with avidity, producing sulphurous acid :—

$$\underset{\text{Hyposulphurous acid.}}{\mathbf{S}Ho_2} + O = \underset{\text{Sulphurous acid.}}{SOHo_2}.$$

SULPHUROUS ANHYDRIDE.

$\mathbf{S}O_2$.

Molecular weight $=64$. *Molecular volume* □□. 1 *litre weighs* 32 *criths*. *Solid at* −76°. *Liquid under the pressure of two atmospheres at* 7°. *Boils at* −8°.

Occurrence.—1. As a volcanic product.

2. In the air of towns.

3. Evolved in the roasting of copper pyrites and other sulphureous ores.

Preparation.—1. By the combustion of sulphur in air or in oxygen:—

$$\mathbf{S} + O_2 = \mathbf{S}O_2.$$

2. By heating sulphuric acid with copper or mercury:—

$$\underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} + Cu = \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + \underset{\text{Cupric sulphate.}}{\mathbf{S}O_2Cuo''} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

$$\underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} + Hg = \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + \underset{\text{Mercuric sulphate.}}{\mathbf{S}O_2Hgo''} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

3. By heating charcoal with sulphuric acid:—

$$\underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} + C = \underset{\text{Sulphurous anhydride.}}{2\mathbf{S}O_2} + \underset{\text{Carbonic anhydride.}}{\mathbf{C}O_2} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

4. By heating a mixture of about three parts by weight of sulphur (two atoms) with four of manganic oxide (one molecule):—

$$S_2 + \underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} = \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + \underset{\text{Manganous sulphide.}}{\mathbf{Mn}S''}.$$

Reactions.—1. Dissolved by water, producing an acid liquid which, when cooled to 0°, deposits white cubical crystals of sulphurous acid:—

$$\underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Sulphurous acid.}}{\mathbf{S}OHo_2}.$$

2. Sulphurous anhydride, when passed into solutions of the metallic hydrates, produces sulphites. If the sulphurous anhydride be in excess, an acid sulphite is obtained:—

$$\underset{\text{Potassic hydrate.}}{\mathbf{O}KH} + \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} = \underset{\text{Hydric potassic sulphite.}}{\mathbf{S}OHoKo}.$$

3. If the metallic hydrate be in excess, a normal sulphite is formed, thus:—

$$\underset{\text{Potassic hydrate.}}{2\mathbf{O}KH} + \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} = \underset{\text{Normal potassic sulphite.}}{\mathbf{S}OKo_2} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

4. Sulphurous acid, when acted upon by metallic hydrates, produces the same salts:—

$$\mathbf{O}KH + \mathbf{S}OHo_2 = \mathbf{S}OHoKo + \mathbf{O}H_2:$$

$$2\mathbf{O}KH + \mathbf{S}OHo_2 = \mathbf{S}OKo_2 + 2\mathbf{O}H_2$$

5. Sulphurous anhydride, when passed over metallic peroxides, produces sulphates :—

$$\underset{\text{Plumbic peroxide.}}{\mathbf{Pb}O_2} + \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} = \underset{\text{Plumbic sulphate.}}{\mathbf{S}O_2Pbo''}.$$

Detection.—Sulphites are recognized by the pungent odour of sulphurous anhydride which they evolve on the addition of a strong acid, such as sulphuric acid :—

$$\underset{\substack{\text{Potassic}\\\text{sulphite.}}}{\mathbf{S}OKo_2} + \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{\mathbf{S}O_2Ho_2} = \underset{\substack{\text{Potassic}\\\text{sulphate.}}}{\mathbf{S}O_2Ko_2} + \underset{\substack{\text{Sulphurous}\\\text{anhydride.}}}{\mathbf{S}O_2} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

When solutions of sulphites are mixed with solution of argentic nitrate, a white precipitate of argentic sulphite is formed :—

$$\underset{\substack{\text{Potassic}\\\text{sulphite.}}}{\mathbf{S}OKo} + \underset{\substack{\text{Argentic}\\\text{nitrate.}}}{2\mathbf{N}O_2Ago} = \underset{\substack{\text{Argentic}\\\text{sulphite.}}}{\mathbf{S}OAgo_2} + \underset{\substack{\text{Potassic}\\\text{nitrate.}}}{2\mathbf{N}O_2Ko}.$$

When this argentic sulphite is boiled with water, it becomes black, owing to the separation of metallic silver :—

$$\underset{\text{Argentic sulphite.}}{\mathbf{S}OAgo_2} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + Ag_2.$$

SULPHURIC ANHYDRIDE.

$\mathbf{S}O_3$.

Molecular weight =80. *Molecular volume* $\square\square$. 1 *litre of sulphuric anhydride vapour weighs* 40 *criths. Fuses at* 24°·5. *Boils at* 52°·6. *Sp. gr.* 1·95.

Preparation.—1. By passing a mixture of sulphurous anhydride and oxygen over ignited spongy platinum :—

$$\underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + O = \underset{\text{Sulphuric anhydride.}}{\mathbf{S}O_3}.$$

2. By heating Nordhausen sulphuric acid :—

$$\left\{\begin{matrix}\mathbf{S}O_2Ho\\ O\\ \mathbf{S}O_2Ho\end{matrix}\right. = \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + \underset{\text{Sulphuric}}{\mathbf{S}O_3}.$$

3. By heating the so-called anhydrous sodic bisulphate (disodic disulphate):—

$$\underset{\substack{\text{Anhydrous sodic bisulphate}\\\text{(Disodic disulphate).}}}{\left\{\begin{matrix}SO_2Nao\\O\\SO_2Nao\end{matrix}\right.} = \underset{\text{Sodic sulphate.}}{SO_2Nao_2} + \underset{\substack{\text{Sulphuric}\\\text{anhydride.}}}{SO_3}.$$

4. By heating sulphuric acid with phosphoric anhydride:—

$$\underset{\substack{\text{Sulphuric}\\\text{acid.}}}{SO_2Ho_2} + \underset{\substack{\text{Phosphoric}\\\text{anhydride.}}}{P_2O_5} = \underset{\substack{\text{Sulphuric}\\\text{anhydride.}}}{SO_3} + \underset{\substack{\text{Meta-phosphoric}\\\text{acid.}}}{2PO_2Ho}.$$

SULPHURIC ACID.

SO_2Ho_2.

Molecular weight =98. *Molecular volume* ⊞. *Dissociation.*

1 litre of sulphuric acid vapour weighs 24·5 *criths. Sp. gr.* 1·85. *Boils at* 325°.

Preparation.—1. By the spontaneous oxidation of sulphur in the presence of water:—

$$S_2 + 3O_2 + \underset{\text{Water.}}{2OH_2} = \underset{\text{Sulphuric acid.}}{2SO_2Ho_2}$$

2. By the action of hydroxyl upon sulphurous anhydride:—

$$\underset{\text{Sulphurous anhydride.}}{SO_2} + \underset{\text{Hydroxyl.}}{Ho_2} = \underset{\text{Sulphuric acid.}}{SO_2Ho_2}.$$

3. By the exposure of a solution of sulphurous acid to air or oxygen:—

$$\underset{\text{Sulphurous acid.}}{SOHo_2} + O = \underset{\text{Sulphuric acid.}}{SO_2Ho_2}.$$

4. By the addition of water to sulphuric anhydride:—

$$\underset{\substack{\text{Sulphuric}\\\text{anhydride.}}}{SO_3} + \underset{\text{Water.}}{OH_2} = \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{SO_2Ho_2}.$$

5. By the action of nitric peroxide and oxygen on sulphurous anhydride and subsequent decomposition by water of the white

crystalline compound thus produced (Brüning and De la Provostaye):—

$$\underset{\text{Sulphurous anhydride.}}{2\mathbf{S}O_2} + \underset{\text{Nitric peroxide.}}{\mathbf{N}^{\text{iv}}{}_2O_4} + O = \underset{\text{White crystalline compound.}}{\left\{\begin{array}{l}\mathbf{S}O_2(N^{\text{v}}O_2)\\ O\\ \mathbf{S}O_2(N^{\text{v}}O_2)\end{array}\right.}.$$

$$\underset{\text{White crystalline compound*.}}{\left\{\begin{array}{l}\mathbf{S}O_2(N^{\text{v}}O_2)\\ O\\ \mathbf{S}O_2(N^{\text{v}}O_2)\end{array}\right.} + \underset{\text{Water.}}{2\mathbf{O}H_2} = \underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} + \underset{\text{Nitrous anhydride.}}{\mathbf{N}_2O_3}.$$

In the manufacture of sulphuric acid on the large scale, the nitrous anhydride is again acted on by water and transformed into nitric acid and nitric oxide:—

$$\underset{\text{Nitrous anhydride.}}{3\mathbf{N}_2O_3} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} + \underset{\text{Nitric oxide.}}{2\mathbf{N}''{}_2O_2}.$$

The nitric oxide, by the action of oxygen, reproduces nitric peroxide, which is then ready to undergo the same processes a second time. The nitric acid is at the same time reduced to nitric peroxide by the action of sulphurous anhydride:—

$$\underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} = \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + \underset{\text{Nitric peroxide.}}{\mathbf{N}^{\text{iv}}{}_2O_4}.$$

The crude sulphuric acid may be freed from traces of nitrous anhydride (which it always contains) by the addition of some ammonic sulphate:—

$$\underset{\text{Ammonic sulphate.}}{\mathbf{S}O_2(NH_4O)_2} + \underset{\text{Nitrous anhydride.}}{\mathbf{N}_2O_3} = \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + \underset{\text{Water.}}{3\mathbf{O}H_2} + N_4.$$

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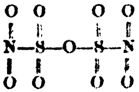
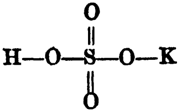
* 

$$\begin{array}{ccccccccc} O & & O & & & & O & & O\\ \| & & \| & & & & \| & & \|\\ N & — & S & — & O & — & S & — & N\\ \| & & \| & & & & \| & & \|\\ O & & O & & & & O & & O \end{array}$$

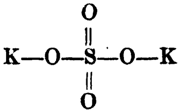
Character.—Sulphuric acid forms several classes of salts:—

Hydric potassic sulphate .... $\mathbf{S}O_2KoHo.$



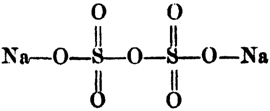
$$\mathrm{H{-}O{-}}\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}\mathrm{{-}O{-}K}$$

Potassic sulphate... $\mathbf{S}O_2Ko_2.$



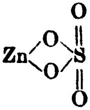
$$\mathrm{K{-}O{-}}\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}\mathrm{{-}O{-}K}$$

Anhydrous sodic bisulphate. (*Disodic disulphate.*) ...... $\left\{\begin{array}{l}\mathbf{S}O_2Nao\\ O\\ \mathbf{S}O_2Nao\end{array}\right.$.

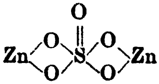


$$\mathrm{Na{-}O{-}}\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}\mathrm{{-}O{-}}\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}\mathrm{{-}O{-}Na}$$

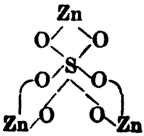
Zincic sulphate ...... $\mathbf{S}O_2Zno''.$



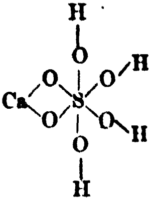
Tetrabasic zincic sulphate. (*Dizincic sulphate.*) .... $\mathbf{S}OZno''_2.$



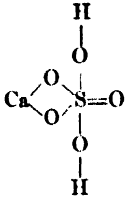
Hexabasic zincic sulphate. (*Trizincic sulphate.*) .... $\mathbf{S}Zno''_3.$



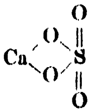
Crystallized gypsum. (*Tetrahydric calcic sulphate.*) ... } $\mathbf{S}Ho_4Cao''$.



Gypsum dried at 100°. (*Dihydric calcic sulphate.*) ... } $\mathbf{S}OHo_2Cao''$.



Gypsum dried at 260°. (*Calcic sulphate.*) .... } $\mathbf{S}O_2Cao''$.



THIOSULPHURIC ACID, *Hyposulphurous Acid.*

$\mathbf{S}S''OHo_2$ (hypothetical).

Preparation of Thiosulphates.—1. By boiling a solution of sodic sulphite with sulphur:—

$$\mathbf{S}ONao_2 \quad + \quad S \quad = \quad \mathbf{S}S''ONao_2.$$

Sodic sulphite. Sodic thiosulphate.

2. By exposure of an alkaline persulphide to the air:—

$$'\mathbf{S}'_2Ca'' \quad + \quad O_2 \quad = \quad \mathbf{S}S''OCao''.$$

Calcic persulphide. Calcic thiosulphate.

Reaction.—The thiosulphates, when acted upon by acids, evolve sulphurous anhydride, whilst sulphur is precipitated:—

$$\mathbf{S}S''ONao_2 + 2HCl = 2NaCl + \mathbf{O}H_2 + S + \mathbf{S}O_2$$

Sodic thiosulphate. Hydrochloric acid. Sodic chloride. Water. Sulphurous anhydride.

DITHIONIC ACID, *Hyposulphuric Acid.*

$'\mathbf{S}^{v}{}_2O_4Ho_2.$

Preparation.—Powdered manganic oxide is suspended in water and a current of sulphurous anhydride passed through the liquid, when the manganic oxide gradually dissolves. The solution contains manganous dithionate or hyposulphate:—

$$\underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} + \underset{\text{Sulphurous anhydride.}}{2\mathbf{S}O_2} = \underset{\text{Manganous dithionate.}}{'\mathbf{S}^{v}{}_2O_4Mno''}.$$

This solution is next treated with baric sulphide, which precipitates manganous sulphide, baric dithionate existing in the solution:—

$$\underset{\text{Manganous dithionate.}}{'\mathbf{S}^{v}{}_2O_4Mno''} + \underset{\text{Baric sulphide.}}{\mathbf{Ba}S''} = \underset{\text{Manganous sulphide.}}{\mathbf{Mn}S''} + \underset{\text{Baric dithionate.}}{'\mathbf{S}^{v}{}_2O_4Bao''}.$$

By adding sulphuric acid to a solution of the baric dithionate, baric sulphate is precipitated and dithionic acid remains in solution:—

$$\underset{\text{Baric dithionate.}}{'\mathbf{S}^{v}{}_2O_4Bao''} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Baric sulphate.}}{\mathbf{S}O_2Bao''} + \underset{\text{Dithionic acid.}}{'\mathbf{S}^{v}{}_2O_4Ho_2}.$$

TRITHIONIC ACID, *Sulphodithionic Acid, Sulphuretted Hyposulphuric Acid.*

$$\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ \mathbf{S}O_2Ho\end{array}\right..$$

Preparation.—By digesting hydric potassic sulphite with sulphur, potassic trithionate and potassic thiosulphate (*hyposulphite*) are formed:—

$$\underset{\text{Hydric potassic sulphite.}}{6\mathbf{S}OKoHo} + 2S = \underset{\text{Potassic trithionate.}}{2\left\{\begin{array}{l}\mathbf{S}O_2Ko\\ S''\\ \mathbf{S}O_2Ko\end{array}\right.} + \underset{\text{Potassic thiosulphate.}}{\mathbf{S}S''OKo_2} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

The two salts so produced, when decomposed by hydrofluosilicic acid, yield trithionic acid, sulphurous acid, and sulphur:—

$$\underset{\text{Potassic thiosulphate.}}{\mathbf{S}S''OKo_2} + 2\underset{\text{Potassic trithionate.}}{\left\{\begin{array}{l}\mathbf{S}O_2Ko\\ S''\\ \mathbf{S}O_2Ko\end{array}\right.} + \underset{\text{Hydrofluosilicic acid.}}{3H_2Si^{iv}F_6} = \underset{\text{Potassic silicofluoride.}}{3K_2Si^{iv}F_6}$$

$$+ 2\underset{\text{Trithionic acid.}}{\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ \mathbf{S}O_2Ho\end{array}\right.} + \underset{\text{Sulphurous acid.}}{\mathbf{S}OHo_2} + S.$$

TETRATHIONIC ACID, *Disulphodithionic Acid, Bisulphuretted Hyposulphuric Acid.*

$$\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ S''\\ \mathbf{S}O_2Ho\end{array}\right. .$$

Preparation.—When iodine is added to baric thiosulphate (*hyposulphite*), baric iodide and baric tetrathionate are produced:—

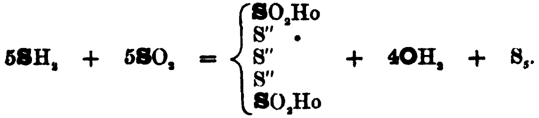
$$\underset{\text{Baric thiosulphate.}}{2\mathbf{S}S''OBao''} + I_2 = \underset{\text{Baric iodide.}}{\mathbf{Ba}I_2} + \underset{\text{Baric tetrathionate.}}{\left\{\begin{array}{l}\mathbf{S}O_2\urcorner\\ S''\\ S''\ \ Bao''\\ \mathbf{S}O_2\lrcorner\end{array}\right.} .$$

This salt, when decomposed by sulphuric acid, yields tetrathionic acid.

PENTATHIONIC ACID, *Trisulphodithionic Acid, Trisulphuretted Hyposulphuric Acid.*

$$\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ S''\\ S''\\ S''\\ \mathbf{S}O_2He\end{array}\right. .$$

Preparation.—This acid is obtained by the action of hydrosulphuric acid on sulphurous anhydride :—



$$5\mathbf{S}H_2 + 5\mathbf{S}O_2 = \begin{cases} \mathbf{S}O_2Ho \\ S'' \\ S'' \\ S'' \\ \mathbf{S}O_2Ho \end{cases} + 4\mathbf{O}H_2 + S_5.$$

Sulphuretted hydrogen. Sulphurous anhydride. Pentathionic acid. Water.

SELENIUM, Se_2.

Atomic weight =79. *Molecular weight* =158. *Molecular volume* □□. 1 *litre of selenium vapour weighs* 79 *criths. Sp. gr.* 4·8. *Fuses a little above* 100°. *Boils at about* 700°. *Atomicity* $''$, ^{iv}, *and* ^{vi}. *Evidence of atomicity* :—

Hydroselenic acid	$\mathbf{Se}''H_2$.
Selenious chloride	$\mathbf{Se}^{iv}Cl_4$.
Selenic acid	$\mathbf{Se}^{vi}O_2Ho_2$.

Occurrence.—In small quantities in some mineral sulphides.

COMPOUNDS OF SELENIUM WITH HYDROGEN AND CHLORINE.

SELENIURETTED HYDROGEN, *Hydroselenic Acid.*

$\mathbf{Se}H_2$.

Molecular weight =81. *Molecular volume* □□. 1 *litre weighs* 40·5 *criths.*

Preparation.—By the action of hydrochloric acid upon ferrous selenide :—

$$\mathbf{Fe}Se'' + 2HCl = \mathbf{Se}H_2 + \mathbf{Fe}Cl_2.$$

Ferrous selenide. Hydrochloric acid. Hydroselenic acid. Ferrous chloride.

Character.—Like hydrosulphuric acid, it produces precipitates in solutions of most of the heavy metals.

There are two chlorides of selenium : $'\mathbf{Se}'_2Cl$ and $'\mathbf{Se}Cl_4$.

COMPOUNDS OF SELENIUM WITH OXYGEN AND HYDROXYL.

Selenious anhydride	$\mathbf{Se}O_2$.
Selenious acid	$\mathbf{Se}OHo_2$.
Selenic acid	$\mathbf{Se}O_2Ho_2$.

These bodies closely resemble the corresponding sulphur compounds.

Selenious anhydride is formed by burning selenium in oxygen:—

$$Se_2 \quad + \quad 2O_2 \quad = \quad \underset{\text{Selenious anhydride.}}{2\mathbf{Se}O_2.}$$

Selenious acid is formed by dissolving the anhydride in boiling water and crystallizing.

Potassic seleniate is prepared by fusing selenium or metallic selenides with nitre. The acid is obtained by transforming the potassic salt into a plumbic salt, and subsequently decomposing the latter with hydrosulphuric acid.

TELLURIUM, Te_2.

Atomic weight = 128. *Molecular weight* = 256. *Sp. gr.* 6·2. *Fuses at* 490°–500°. *Atomicity* $''$, ^{iv}, *and* ^{vi}. *Evidence of atomicity*:—

Hydrotelluric acid	$\mathbf{Te}''H_2$.
Tellurous chloride	$\mathbf{Te}^{iv}Cl_4$.
Telluric acid	$\mathbf{Te}^{vi}O_2Ho_2$.

This element is of even less importance than selenium, which it closely resembles.

The following compounds are known :—

Hydrotelluric acid (*telluretted hydrogen*)	**TeH_2.**
Hypotellurous chloride	**$TeCl_2$.**
Tellurous chloride	**$TeCl_4$.**
Tellurous anhydride	**TeO_2.**
Telluric anhydride	**TeO_3.**
Tellurous acid	**$TeOHo_2$?**
Telluric acid	**TeO_2Ho_2.**

CHAPTER XIV.

MONAD ELEMENTS.

SECTION II. (*continued from* Chap. VIII.).

BROMINE, Br_2.

Atomic weight $=80$. *Molecular weight* $=160$. *Molecular volume* □□. 1 *litre of bromine vapour weighs* 80 *criths. Sp. gr.* 3·18. *Fuses at* $-20°$. *Boils at* $63°$. *Atomicity* '. *Evidence of Atomicity* :—

Hydrobromic acid	HBr.
Potassic bromide	KBr.
Argentic bromide	$AgBr$.

Occurrence.—In small quantities in some saline mineral waters. In sea-water, and the waters of the Dead Sea.

Preparation.—1. By the treatment with chlorine, of the mother-liquors of saline waters containing bromides, and extracting the liberated bromine by ether :—

$$\underset{\text{Potassic bromide.}}{2KBr} + Cl_2 = \underset{\text{Potassic chloride.}}{2KCl} + Br_2.$$

2. By heating together sulphuric acid, sodic bromide, and manganic oxide:—

$$\underset{\text{Sodic bromide.}}{2NaBr} + \underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} + \underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} = Br_2$$

$$+ \underset{\text{Sodic sulphate.}}{\mathbf{S}O_2Nao_2} + \underset{\text{Manganous sulphate.}}{\mathbf{S}O_2Mno''} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

Character.—Bromine unites with several metals directly, and with great energy. Antimony and arsenic burn in it with brilliancy.

At 0° bromine combines with water, forming a crystalline compound, $Br_2, 10\mathbf{O}H_2$.

HYDROBROMIC ACID.

HBr.

Molecular weight =81. *Molecular volume* ☐☐. 1 *litre of hydrobromic acid weighs* 40·5 *criths. Fuses at* −73°. *Boils at* −69°.

Preparation.—1. By passing a mixture of hydrogen and bromine vapour through a red-hot tube, or by burning hydrogen in a mixture of bromine vapour and air:—

$$H_2 + Br_2 = \underset{\text{Hydrobromic acid.}}{2HBr}.$$

2. By heating potassic bromide with phosphoric acid:—

$$\underset{\text{Potassic bromide.}}{3KBr} + \underset{\text{Phosphoric acid.}}{\mathbf{P}OHo_3} = \underset{\text{Potassic phosphate.}}{\mathbf{P}OKo_3} + \underset{\text{Hydrobromic acid.}}{3HBr}.$$

Sulphuric acid cannot be employed for this operation, as a portion of the hydrobromic acid is then decomposed, bromine being liberated:—

$$\underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} + \underset{\text{Hydrobromic acid.}}{2HBr} = Br_2 + \underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2}.$$

3. By the action of water upon phosphorous tribromide:—

$$\underset{\text{Phosphorous tribromide.}}{\mathbf{P}'''Br_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\text{Phosphorous acid.}}{\mathbf{P}{\cdot}OHHo_2} + \underset{\text{Hydrobromic acid.}}{3HBr}.$$

4. By gradually dropping bromine into water containing amorphous phosphorus :—

$$P_2 + 3Br_2 + \underset{\text{Water.}}{6\mathbf{O}H_2} = \underset{\text{Phosphorous acid.}}{2\mathbf{P}^{v}OHHo_2} + \underset{\text{Hydrobromic acid.}}{6HBr.}$$

5. By passing sulphuretted hydrogen through water containing bromine :—

$$\underset{\text{Sulphuretted hydrogen.}}{2\mathbf{S}H_2} + 2Br_2 = \underset{\text{Hydrobromic acid.}}{4HBr} + S_2.$$

Reactions.—1. Decomposed by chlorine with liberation of bromine :—

$$\underset{\text{Hydrobromic acid.}}{2HBr} + Cl_2 = \underset{\text{Hydrochloric acid.}}{2HCl} + Br_2.$$

2. By the action of atmospheric oxygen a small quantity of bromine is liberated ; but the decomposition is soon arrested :—

$$\underset{\text{Hydrobromic acid.}}{4HBr} + O_2 = \underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Bromine.}}{2Br_2.}$$

3. In contact with metallic oxides, hydrates, and salts, bromides are formed.

COMPOUNDS OF BROMINE WITH OXYGEN AND HYDROXYL.

Hypobromous anhydride	$\mathbf{O}Br_2.$
Hypobromous acid ……	$\mathbf{O}BrH.$
Bromic acid ……………	$\left\{\begin{array}{l}\mathbf{O}Br\\ O\ \ .\\ \mathbf{O}H\end{array}\right.$

The graphic formulæ of these compounds are analogous to those of the corresponding chlorine compounds, given at page 46.

HYPOBROMOUS ANHYDRIDE.

$\mathbf{O}Br_2$.

Preparation.—By passing bromine vapour over dry mercuric oxide :—

$$\underset{\text{Mercuric oxide.}}{\mathbf{Hg}O} + 2Br_2 = \underset{\text{Mercuric bromide.}}{\mathbf{Hg}Br_2} + \underset{\text{Hypobromous anhydride.}}{\mathbf{O}Br_2}.$$

HYPOBROMOUS ACID.

$\mathbf{O}BrH$.

Preparation.—1. By passing hypobromous anhydride into water :—

$$\underset{\text{Hypobromous anhydride.}}{\mathbf{O}Br_2} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Hypobromous acid.}}{2\mathbf{O}BrH}.$$

2. By agitating mercuric oxide with bromine-water :—

$$\underset{\text{Mercuric oxide.}}{2\mathbf{Hg}O} + \underset{\text{Water.}}{\mathbf{O}H_2} + 2Br_2 = \underset{\text{Hypobromous acid.}}{2\mathbf{O}BrH} + \underset{\text{Mercuric oxybromide.}}{\left\{\begin{array}{l}\mathbf{Hg}Br\\ O\\ \mathbf{Hg}Br\end{array}\right.}.$$

BROMIC ACID.

$$\left\{\begin{array}{l}\mathbf{O}Br\\ \mathbf{O}Ho\end{array}\right.$$

Preparation.—By acting upon a solution of baric bromate with sulphuric acid :—

$$\underset{\text{Baric bromate.}}{\left\{\begin{array}{l}\mathbf{O}Br\\ O\\ Bao''\\ O\\ \mathbf{O}Br\end{array}\right.} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Bromic acid.}}{2\left\{\begin{array}{l}\mathbf{O}Br\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Baric sulphate.}}{\mathbf{S}O_2Bao''}.$$

Reaction.—By boiling, bromic acid decomposes into water, bromine, and oxygen:—

$$4\underset{\text{Bromic acid.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Br}\\ \mathbf{O}\mathrm{Ho}\end{matrix}\right.} = 2\mathrm{Br}_2 + \underset{\text{Water.}}{2\mathbf{O}\mathrm{H}_2} + 5\mathrm{O}_2.$$

Preparation of bromates.—1. By adding bromine to a solution of a metallic hydrate, and separating the bromate by crystallization:—

$$\underset{\text{Potassic hydrate.}}{6\mathrm{KHo}} + 3\mathrm{Br}_2 = \underset{\text{Potassic bromide.}}{5\mathrm{KBr}} + \underset{\text{Potassic bromate.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Br}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right.} + \underset{\text{Water.}}{3\mathbf{O}\mathrm{H}_2}.$$

2. By the action of potassic hydrate on bromine pentachloride:—

$$\underset{\text{Potassic hydrate.}}{6\mathrm{KHo}} + \underset{\text{Bromine pentachloride.}}{\mathrm{BrCl}_5} = \underset{\text{Potassic chloride.}}{5\mathrm{KCl}} + \underset{\text{Potassic bromate.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Br}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right.} + \underset{\text{Water.}}{3\mathbf{O}\mathrm{H}_2}.$$

Character of bromates.—Some of the bromates when heated lose oxygen, being transformed into bromides:—

$$2\underset{\text{Potassic bromate.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Br}\\ \mathbf{O}\mathrm{Ko}\end{matrix}\right.} = \underset{\text{Potassic bromide.}}{2\mathrm{KBr}} + 3\mathrm{O}_2.$$

Others evolve bromine and a portion of their oxygen, leaving metallic oxides:—

$$2\underset{\text{Magnesic bromate.}}{\left\{\begin{matrix}\mathbf{O}\mathrm{Br}\\ \mathrm{O}\\ \mathrm{Mgo}''\\ \mathrm{O}\\ \mathbf{O}\mathrm{Br}\end{matrix}\right.} = \underset{\text{Magnesic oxide.}}{2\mathrm{MgO}} + 2\mathrm{Br}_2 + 5\mathrm{O}_2.$$

IODINE, I_2.

Atomic weight =127. ***Molecular weight*** =254. ***Molecular volume*** ☐☐. 1 ***litre of iodine vapour weighs*** 127 ***criths.*** ***Sp. gr.*** 4·95. ***Fuses at*** 107°. ***Boils at*** 180°. ***Atomicity*** ′. ***Evidence of atomicity*** :—

Hydriodic acid	HI.
Potassic iodide	KI.
Argentic iodide	AgI.

Occurrence.—In mineral springs, in sea-water, and in considerable quantities in sea-plants.

Manufacture.—Sea-weeds are burnt or carbonized in iron retorts, and the ash or charcoal is extracted with water. The liquid is evaporated, and, after a considerable quantity of sodic carbonate and chloride has crystallized out, the mother-liquor, which contains potassic iodide, is distilled with sulphuric acid and manganic oxide:—

$$\underset{\text{Potassic iodide.}}{2KI} + \underset{\text{Manganic oxide.}}{\mathbf{Mn}O_2} + \underset{\text{Sulphuric acid.}}{2\mathbf{S}O_2Ho_2} = \underset{\text{Potassic sulphate.}}{\mathbf{S}O_2Ko_2} + \underset{\text{Manganous sulphate.}}{\mathbf{S}O_2Mno''} + I_2 + \underset{\text{Water}}{2\mathbf{O}H_2}.$$

Reactions.—1. Iodine is separated from nearly all its compounds by chlorine and bromine:—

$$\underset{\text{Potassic iodide.}}{2KI} + Cl_2 = \underset{\text{Potassic chlorate.}}{2KCl} + I_2.$$

$$\underset{\text{Potassic iodide.}}{2KI} + Br_2 = \underset{\text{Potassic bromide.}}{2KBr} + I_2.$$

2. Iodine unites directly with many metals.

HYDRIODIC ACID.

HI.

Molecular weight $=128$. *Molecular volume* ☐☐. 1 *litre of hydriodic acid weighs* 64 *criths. Fuses at* $-55°$.

Preparation.—1. By passing iodine vapour and hydrogen through a red-hot tube or over spongy platinum gently heated:—

$$H_2 + I_2 = 2HI.$$

2. By the action of dilute sulphuric acid on baric iodide, or of phosphoric acid on any iodide :—

$$\underset{\text{Baric iodide.}}{\mathbf{Ba}I_2} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Hydriodic acid.}}{2HI} + \underset{\text{Baric sulphate.}}{\mathbf{S}O_2Bao''}.$$

3. By decomposing phosphorous triiodide by water :—

$$\underset{\text{Phosphorous triiodide.}}{\mathbf{P}I_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\text{Phosphorous acid.}}{\mathbf{P}OHHo_2} + \underset{\text{Hydriodic acid.}}{3HI}.$$

4. By heating together water, potassic iodide, iodine, and phosphorus :—

$$\underset{\text{Potassic iodide.}}{4KI} + P_2 + 5I_2 + \underset{\text{Water.}}{8\mathbf{O}H_2} = \underset{\text{Hydriodic acid.}}{14HI} + \underset{\text{Hydric dipotassic phosphate.}}{2\mathbf{P}OHoKo_2}.$$

5. A solution of hydriodic acid is obtained by passing sulphuretted hydrogen through water in which iodine is suspended:—

$$\underset{\text{Sulphuretted hydrogen.}}{2\mathbf{S}H_2} + 2I_2 = \underset{\text{Hydriodic acid.}}{4HI} + S_2.$$

Reactions.—1. Decomposed by chlorine and bromine, with liberation of iodine :—

$$\underset{\text{Hydriodic acid.}}{2HI} + Cl_2 = \underset{\text{Hydrochloric acid.}}{2HCl} + I_2.$$

$$\underset{\text{Hydriodic acid.}}{2HI} + Br_2 = \underset{\text{Hydrobromic acid.}}{2HBr} + I_2.$$

2. It is gradually but completely decomposed by atmospheric oxygen ; the iodone, which at first remains dissolved in the hydriodic acid, is after a time deposited in crystals :—

$$\underset{\text{Hydriodic acid.}}{4HI} + O_2 = \underset{\text{Water.}}{2\mathbf{O}H_2} + 2I_2.$$

3. With metallic oxides, hydrates, and some salts it forms iodides. Even argentic chloride is transformed by hydriodic acid into argentic iodide :—

$$\underset{\text{Argentic chloride.}}{AgCl} + \underset{\text{Hydriodic acid.}}{HI} = \underset{\text{Argentic iodide.}}{AgI} + \underset{\text{Hydrochloric acid.}}{HCl}.$$

4. Hydriodic acid is rapidly decomposed by mercury, with liberation of hydrogen:—

$$\underset{\substack{\text{Hydriodic}\\ \text{acid.}}}{2HI} + 2Hg = \underset{\substack{\text{Mercurous}\\ \text{iodide.}}}{{}'Hg'_2I_2} + H_2.$$

COMPOUNDS OF IODINE WITH OXYGEN AND HYDROXYL.

$$\text{Iodic anhydride} \ldots\ldots\ldots\ldots \left\{\begin{array}{l}\mathbf{O}I\\ O\\ O\\ O\\ \mathbf{O}I.\end{array}\right.$$

$$\text{Iodic acid} \ldots\ldots\ldots\ldots\ldots\ldots \left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right.$$

$$\text{Periodic anhydride} \ldots\ldots\ldots \left\{\begin{array}{l}\mathbf{O}I\\ O\\ O\\ O\\ O\\ O\\ \mathbf{O}I\end{array}\right.$$

$$\text{Periodic acid} \ldots\ldots\ldots\ldots\ldots \left\{\begin{array}{l}\mathbf{O}I\\ O\\ \mathbf{O}Ho\end{array}\right. .$$

The graphic formulæ of these compounds are analogous to those of the corresponding chlorine compounds given at p. 46.

IODIC ANHYDRIDE.

$$I_2O_5, \text{ or } \left\{\begin{array}{l}\mathbf{O}I\\ O\\ O\\ O\\ \mathbf{O}I\end{array}\right. .$$

Preparation.—By heating iodic acid to 170°, when it separates into iodic anhydride and water:—

$$\underset{\text{Iodic acid.}}{2\left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right.} = \mathbf{O}H_2 + \underset{\text{Iodic anhydride.}}{\left\{\begin{array}{l}\mathbf{O}I\\ O\\ O\\ O\\ \mathbf{O}I\end{array}\right.}$$

Reaction.—When strongly heated, it decomposes into iodine and oxygen.

IODIC ACID.

$$\left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right..$$

Preparation.—1. By the action of sulphuric acid upon baric iodate:—

$$\underset{\text{Baric iodate.}}{\left\{\begin{array}{l}\mathbf{O}I\\ O\\ Bao''\\ O\\ \mathbf{O}I\end{array}\right.} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Iodic acid.}}{2\left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Baric sulphate.}}{\mathbf{S}O_2Bao''}.$$

2. By oxidizing iodine with strong boiling nitric acid:—

$$\underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho} + I_2 = \underset{\text{Iodic acid.}}{2\left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Nitrous anhydride.}}{2\mathbf{N}_2O_3} + \underset{\text{Nitric peroxide.}}{'\mathbf{N}^{\text{iv}}{}_2O_4}.$$

3. By acting upon iodine and water with chlorine:—

$$I_2 + \underset{\text{Water.}}{6\mathbf{O}H_2} + 5Cl_2 = \underset{\text{Iodic acid.}}{2\left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Hydrochloric acid.}}{10HCl}.$$

Reactions.—1. In contact with hydriodic acid it forms water and iodine:—

$$\underset{\text{Iodic acid.}}{\left\{\begin{array}{l}\mathbf{O}I\\ \mathbf{O}Ho\end{array}\right.} + \underset{\text{Hydriodic acid.}}{5HI} = \underset{\text{Water.}}{3\mathbf{O}H_2} + 3I_2.$$

2. It is reduced by many other deoxidizing agents.

Preparation of Iodates.—1. By treating solutions of metallic hydrates with iodine, and separating the iodate by crystallization :—

$$\underset{\text{Potassic hydrate.}}{6KHo} + 3I_2 = \underset{\text{Potassic iodide.}}{5KI} + \underset{\text{Potassic iodate.}}{\left\{\begin{matrix}OI\\OKo\end{matrix}\right.} + \underset{\text{Water.}}{3OH_2}.$$

2. By dissolving iodine in potassic hydrate and treating the mixture with chlorine :—

$$\underset{\text{Potassic hydrate.}}{12KHo} + I_2 + 5Cl_2 = \underset{\text{Potassic chloride.}}{10KCl} + \underset{\text{Potassic iodate.}}{2\left\{\begin{matrix}OI\\OKo\end{matrix}\right.} + \underset{\text{Water.}}{6OH_2}.$$

3. By heating together potassic chlorate and iodine :—

$$I_2 + \underset{\text{Potassic chlorate.}}{\left\{\begin{matrix}OCl\\OKo\end{matrix}\right.} = \underset{\text{Iodine monochloride.}}{ICl} + \underset{\text{Potassic iodate.}}{\left\{\begin{matrix}OI\\OKo\end{matrix}\right.}.$$

Character of Iodates.—Some of the iodates when heated split into iodides and oxygen, others into metallic oxides, iodine, and oxygen.

Iodic acid gives several well-defined anhydro-salts.

PERIODIC ANHYDRIDE.

$$I_2O_7, \text{ or } \left\{\begin{matrix}OI\\O\\O\\O\\O\\O\\OI\end{matrix}\right.$$

Preparation.—By heating periodic acid to 160° :—

$$\underset{\text{Periodic acid.}}{2\left\{\begin{matrix}OI\\O\\OHo\end{matrix}\right.} = \underset{\text{Periodic anhydride.}}{I_2O_7} + \underset{\text{Water.}}{OH_2}.$$

Reaction.—When heated it is decomposed into oxygen and iodic anhydride, and ultimately into iodine and oxygen.

PERIODIC ACID.

$$\left\{\begin{matrix}\mathbf{O}I\\ O\\ \mathbf{O}Ho\end{matrix}\right. .$$

Preparation.—By decomposing plumbic periodate with sulphuric acid:—

$$\begin{array}{ccccccc} \left\{\begin{matrix}\mathbf{O}I\\ O\\ O\\ Pbo''\\ O\\ O\\ \mathbf{O}I\end{matrix}\right. & + & \mathbf{S}O_2Ho_2 & = & 2\left\{\begin{matrix}\mathbf{O}I\\ O\\ \mathbf{O}Ho\end{matrix}\right. & + & \mathbf{S}O_2Pbo''. \\ \text{Plumbic periodate.} & & \text{Sulphuric acid.} & & \text{Periodic acid.} & & \text{Plumbic sulphate.} \end{array}$$

Preparation of Periodates.—Sodic periodate may be prepared by passing chlorine through mixed solutions of sodic hydrate and sodic iodate:—

$$\begin{array}{ccccccccc} \left\{\begin{matrix}\mathbf{O}I\\ \mathbf{O}Nao\end{matrix}\right. & + & 2NaHo & + & Cl_2 & = & \left\{\begin{matrix}\mathbf{O}I\\ O\\ \mathbf{O}Nao\end{matrix}\right. & + \ \mathbf{O}H_2 \ + & 2NaCl. \\ \text{Sodic iodate.} & & \text{Sodic hydrate.} & & & & \text{Sodic periodate.} & \text{Water.} & \text{Sodic chloride.} \end{array}$$

FLUORINE, F_2.

Atomic weight =19. *Molecular weight* =38 (?). *Molecular volume* ☐☐. 1 *litre weighs* 19 *criths* (?). *Atomicity* '. *Evidence of atomicity*:—

Hydrofluoric acid.............. HF.

Occurrence.—In combination with metals in fluorspar, cryolite, apatite, and other minerals. Widely but sparsely diffused. Little is known of fluorine in the uncombined condition.

COMPOUND OF FLUORINE WITH HYDROGEN.

HYDROFLUORIC ACID.

HF.

Molecular weight =20. *Molecular volume* □□. 1 *litre weighs* 10 *criths*. *Boils at* −19°·5. *Sp. gr. of liquid* ·9875 *at* 18°.

Preparation.—By heating calcic fluoride with sulphuric acid in a leaden or platinum vessel:—

$$\underset{\substack{\text{Calcic}\\\text{fluoride.}}}{\mathbf{Ca}F_2} + \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{\mathbf{S}O_2Ho_2} = \underset{\substack{\text{Hydrofluoric}\\\text{acid.}}}{2HF} + \underset{\substack{\text{Calcic}\\\text{sulphate.}}}{\mathbf{S}O_2Cao''}.$$

CHAPTER XV.

TETRAD ELEMENTS.

SECTION I. (*Continued from* Chapter XI.)

SILICON, *Silicium*, Si.

Atomic weight =28·5. *Sp. gr.* (*graphitoidal*) =2·49. *Atomicity*$^{\text{iv}}$.

Evidence of atomicity:—

Silicic chloride .... $\mathbf{Si}Cl_4$.
Silicic fluoride.............. $\mathbf{Si}F_4$.

Occurrence.—Silicon is one of the most widely diffused elements. It is found, in combination with oxygen and metals, in a very large number of minerals.

a. Amorphous Silicon.

Preparation.—1. By heating potassic silicofluoride with potassium:—

$$\underset{\substack{\text{Potassic}\\\text{silicofluoride.}}}{SiK_2F_6} + 2K_2 = Si + \underset{\substack{\text{Potassic}\\\text{fluoride.}}}{6KF}.$$

2. By heating sodium in a current of the vapour of silicic chloride :—

$$\underset{\substack{\text{Silicic}\\\text{chloride.}}}{\mathbf{Si}Cl_4} + 2Na_2 = Si + \underset{\substack{\text{Sodic}\\\text{chloride.}}}{4NaCl.}$$

Reactions.—1. Silicon is dissolved by aqueous hydrofluoric acid, and converted into hydrofluosilicic acid :—

$$Si + \underset{\substack{\text{Hydrofluoric}\\\text{acid.}}}{6HF} = \underset{\substack{\text{Hydrofluosilicic}\\\text{acid.}}}{SiH_2F_6} + 2H_2.$$

2. When fused with potassic hydrate, or boiled in its solution, it yields potassic silicate :—

$$Si + \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{4KHo} = \underset{\substack{\text{Potassic}\\\text{silicate.}}}{\mathbf{Si}Ko_4} + 2H_2.$$

3. Heated in the air, it burns, producing silicic anhydride.

β. *Graphitoidal Silicon.*

Preparation.—By fusing amorphous silicon with aluminium, and boiling the compound in hydrochloric or hydrofluoric acid, which dissolves the aluminium, leaving the silicon in the form of hexagonal plates with a metallic lustre.

Character.—May be heated to whiteness in oxygen without burning.

Is gradually oxidized by a mixture of nitric and hydrofluoric acids.

Is slowly attacked by fused potassic hydrate.

γ. *Adamantine Silicon.*

Preparation.—1. By heating aluminium very strongly in a current of the vapour of silicic chloride. The aluminic chloride which is formed volatilizes, leaving the adamantine silicon behind :—

$$\underset{\substack{\text{Silicic}\\\text{chloride.}}}{3\mathbf{Si}Cl_4} + 2Al_2 = \underset{\substack{\text{Aluminic}\\\text{chloride.}}}{2\,{}'\mathbf{Al}'''_2Cl_6} + Si_3.$$

2. By fusing amorphous silicon.

SILICIC HYDRIDE.

$\mathbf{Si}H_4$.

Molecular weight =32·5.

Preparation.—1. By decomposing dilute sulphuric acid by a feeble electric current passing from electrodes of aluminium containing silicon, when the silicic hydride is evolved at the negative pole.

2. By decomposing magnesic silicide with hydrochloric acid:—

$$\underset{\text{Magnesic silicide.}}{\mathbf{Si}Mg''_2} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Magnesic chloride.}}{2\mathbf{Mg}Cl_2} + \underset{\text{Silicic hydride.}}{\mathbf{Si}H_4}.$$

3. Both the above processes furnish silicic hydride mixed with much hydrogen; but if ethylic silicoformate be placed in contact with sodium, it splits up into ethylic silicate and pure silicic hydride, the sodium remaining unaffected:—

$$\underset{\text{Ethylic silicoformate}}{4\mathbf{Si}H(C_2H_5O)_3} = \underset{\text{Silicic hydride.}}{\mathbf{Si}H_4} + \underset{\text{Ethylic silicate.}}{3\mathbf{Si}(C_2H_5O)_4}.$$

Reactions.—1. Under reduced pressure or mixed with hydrogen, inflames spontaneously in air, producing water and silicic anhydride:—

$$\underset{\text{Silicic hydride.}}{\mathbf{Si}H_4} + 2O_2 = \underset{\text{Silicic anhydride.}}{\mathbf{Si}O_2} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

2. Decomposed by solution of potassic hydrate, yielding exactly four times its volume of hydrogen:—

$$\underset{\text{Silicic hydride.}}{\mathbf{Si}H_4} + \underset{\text{Potassic hydrate.}}{2KHo} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Potassic silicate.}}{\mathbf{Si}OKo_2} + 4H_2.$$

SILICIC CHLORIDE.

$\mathbf{Si}Cl_4$.

Molecular weight =170·5. *Molecular volume* $\square\!\square$. 1 *litre weighs* 85·25 *criths.* *Sp. gr. of liquid* 1·52. *Boils at* 59°.

Preparation.—1. By burning silicon in chlorine.

2. By heating a mixture of carbon and silicic anhydride in a stream of chlorine:—

$$\underset{\text{Silicic anhydride.}}{\mathbf{Si}O_2} + 2C + 2Cl_2 = \underset{\text{Silicic chloride.}}{\mathbf{Si}Cl_4} + \underset{\text{Carbonic oxide.}}{2\mathbf{C}O.}$$

Reaction.—By contact with water it produces silicic and hydrochloric acids:—

$$\underset{\text{Silicic chloride.}}{\mathbf{Si}Cl_4} + \underset{\text{Water.}}{4\mathbf{O}H_2} = \underset{\text{Silicic acid.}}{\mathbf{Si}Ho_4} + \underset{\text{Hydrochloric acid.}}{4HCl.}$$

SILICIC HYDROTRICHLORIDE, *Silicon Chloroform.*

$$\mathbf{Si}HCl_3.$$

Molecular weight =136. *Molecular volume* $\square\square$. 1 *litre weighs* 68 *criths. Boils at* 36°.

Preparation.—By heating crystallized silicon to dull redness in a current of hydrochloric acid gas:—

$$Si + \underset{\text{Hydrochloric acid.}}{3HCl} = \underset{\text{Silicic hydrotrichloride.}}{\mathbf{Si}HCl_3} + H_2.$$

Reactions.—1. Is decomposed by chlorine at ordinary temperatures:—

$$\underset{\text{Silicic hydrotrichloride.}}{\mathbf{Si}HCl_3} + Cl_2 = \underset{\text{Silicic chloride.}}{\mathbf{Si}Cl_4} + \underset{\text{Hydrochloric acid.}}{HCl.}$$

2. By contact with water it is transformed into *Disilicic hydrotrioxide*, or *Silicoformic anhydride*:—

$$\underset{\text{Silicic hydrotrichloride.}}{2\mathbf{Si}HCl_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\text{Silicoformic anhydride.}}{\left\{\begin{matrix}\mathbf{Si}HO\\ O\\ \mathbf{Si}HO\end{matrix}\right.} + \underset{\text{Hydrochloric acid.}}{6HCl.}$$

SILICIC BROMIDE.

$SiBr_4$.

Molecular weight =348·5. *Sp. gr.* 2·813 at 0°. *Boils at* 153°.

Preparation.—By the same method as that employed for making the chloride, bromine vapour being substituted for chlorine.

Reaction.—Decomposed by water in the same manner as the chloride.

SILICIC IODIDE.

SiI_4.

Molecular weight =536·5. *Molecular volume* ☐☐. *Fuses at* 120°·5. *Boils in carbonic anhydride at* 290°.

Preparation.—By passing iodine vapour and carbonic anhydride over red-hot silicon.

Reactions.—1. Decomposed by water into silicic and hydriodic acids.

2. By absolute alcohol it is decomposed, with production of silicic anhydride, ethylic iodide, and hydriodic acid:—

$$\underset{\substack{\text{Silicic}\\ \text{iodide.}}}{SiI_4} + \underset{\text{Alcohol.}}{2EtHo} = \underset{\substack{\text{Silicic}\\ \text{anhydride.}}}{SiO_2} + \underset{\substack{\text{Ethylic}\\ \text{iodide.}}}{2EtI} + \underset{\substack{\text{Hydriodic}\\ \text{acid.}}}{2HI}.$$

SILICIC FLUORIDE.

SiF_4.

Molecular weight =104·5. *Molecular volume* ☐☐. 1 *litre weighs* 52·25 *criths*. *Fuses at* −140° C. *Condensable gas.*

Preparation.—By heating together silicic anhydride, calcic fluoride, and sulphuric acid:—

$$\underset{\substack{\text{Silicic}\\ \text{anhydride.}}}{SiO_2} + \underset{\substack{\text{Calcic}\\ \text{fluoride.}}}{2CaF_2} + \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{2SO_2Ho_2} = \underset{\substack{\text{Silicic}\\ \text{fluoride.}}}{SiF_4}$$

$$+ \underset{\text{Dihydric calcic sulphate.}}{2SOHo_2Cao''}.$$

Reaction.—By contact with water it produces silicic and hydrofluosilicic acids :—

$$\underset{\substack{\text{Silicic}\\\text{fluoride.}}}{3\mathbf{Si}F_4} + \underset{\text{Water.}}{4\mathbf{O}H_2} = \underset{\substack{\text{Silicic}\\\text{acid.}}}{\mathbf{Si}Ho_4} + \underset{\substack{\text{Hydrofluosilicic}\\\text{acid.}}}{2SiH_2F_6}.$$

By contact with metallic oxides, hydrates, and salts, hydrofluosilicic acid produces silicofluorides, some of which, as the potassic and baric compounds, are insoluble in water :—

$$\underset{\substack{\text{Hydrofluosilicic}\\\text{acid.}}}{SiH_2F_6} + \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{2KHo} = \underset{\substack{\text{Potassic}\\\text{silicofluoride.}}}{SiK_2F_6} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

COMPOUNDS OF SILICON WITH OXYGEN AND HYDROXYL.

Silicic anhydride	$\mathbf{Si}O_2$.
Silicic acid	$\mathbf{Si}Ho_4$ and $\mathbf{Si}OHo_2$.

Other Modifications of Silicic acid.

$\mathbf{Si}_2O_3Ho_2$
$\mathbf{Si}_3O_5Ho_2$
$\mathbf{Si}_4O_7Ho_2$
$\mathbf{Si}_4O_4Ho_8$
$\mathbf{Si}_6O_{10}Ho_4$
$\mathbf{Si}_6O_8Ho_8$
$\mathbf{Si}_8O_{15}Ho_2$.

SILICIC ANHYDRIDE.

$\mathbf{Si}O_2$.

Molecular weight = 60·5. *Sp. gr.* 2·69.

Occurrence.—In the pure state in many minerals, as quartz, agate, &c.

Preparation.—By heating silicic acid to 100°.

SILICIC ACID.

Tetrabasic... $\mathbf{Si}Ho_4$. Dibasic... $\mathbf{SiO}Ho_2$.

Preparation.—1. By treating a solution of a soluble silicate with hydrochloric acid:—

$$\underset{\text{Sodic silicate.}}{\mathbf{Si}Nao_4} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Silicic acid.}}{\mathbf{Si}Ho_4} + \underset{\text{Sodic chloride}}{4NaCl.}$$

2. By passing a stream of carbonic anhydride through a solution of a soluble silicate:—

$$\underset{\text{Sodic silicate.}}{\mathbf{Si}Nao_4} + \underset{\text{Water.}}{4\mathbf{O}H_2} + \underset{\text{Carbonic anhydride.}}{4\mathbf{CO}_2} = \underset{\text{Silicic acid.}}{\mathbf{Si}Ho_4} + \underset{\text{Hydric sodic carbonate.}}{4\mathbf{C}OHoNao.}$$

A reaction similar to this is the cause of the disintegration of granitic rocks.

3. By passing silicic fluoride through water. (See p. 104.)

4. The bibasic silicic acid is said to be produced by the evaporation *in vacuo* at 16° of a solution of the tetrabasic acid in water.

The acid prepared by the first three of the above processes has probably the formula $\mathbf{Si}Ho_4$; by drying in the air, a compound remains containing

$$Si_6H_8O_{16} = \mathbf{Si}_6O_8Ho_8.$$

This last acid, heated to 100°, loses more water, being transformed into

$$Si_6H_4O_{14} = \mathbf{Si}_6O_{10}Ho_4.$$

5. By the action of water on tetrethylic silicate, a compound is produced containing

$$Si_4H_8O_{12} = \mathbf{Si}_4O_4Ho_8.$$

SILICATES.

The soluble alkaline silicates may be prepared by fusing silicic anhydride, in the form of sand or flints, or insoluble natural silicates, with alkaline hydrates or carbonates.

The silicates form a very important class of minerals. The following list contains a few examples:—

Sand. Flint. Rock crystal. Quartz. Opal. Chalcedony.	*Silicic anhydride.*	$\mathbf{Si}\mathrm{O}_2$.
Peridote	*Dimagnesic silicate*	$\mathbf{Si}\mathrm{Mgo}''_2$.
Phenacite.	*Diglucinic silicate*	$\mathbf{Si}\mathrm{Glo}''_2$.
Willemite.	*Dizincic silicate*	$\mathbf{Si}\mathrm{Zno}''_2$.
Zircon.	*Zirconic silicate*	$\mathbf{Si}\mathrm{Zro}^{\mathrm{iv}}$.
Enstatite.	*Monomagnesic silicate*	$\mathbf{Si}\mathrm{OMgo}''$.
Yorke's Sodic silicate		$\left\{\begin{array}{l}\mathbf{Si}\mathrm{Nao}_3\\ \mathrm{O}\\ \mathbf{Si}\mathrm{Nao}_3\end{array}\right.$.
Ophite (Noble Serpentine)		$\left\{\begin{array}{l}\mathbf{Si}\text{——}\urcorner\\ \mathrm{O}\quad\ \mathrm{Mgo}''_3.\\ \mathbf{Si}\text{——}\lrcorner\end{array}\right.$
Diopside.	*Calcic magnesic disilicate*	$\begin{array}{l}\mathbf{Si}\mathrm{O}\\ \mathbf{Si}\mathrm{O}\end{array}\mathrm{Cao}''\mathrm{Mgo}''$.
Talc.	*Tetramagnesic pentasilicate*	$\mathbf{Si}_5\mathrm{O}_6\mathrm{Mgo}''_4$.
Okenite.	*Tetrahydric calcic disilicate*	$\left\{\begin{array}{l}\mathbf{Si}\mathrm{Ho}_2\text{—}\urcorner\\ \mathrm{O}\quad\ \mathrm{Cao}''.\\ \mathbf{Si}\mathrm{Ho}_2\text{—}\lrcorner\end{array}\right.$

Serpentine. *Dihydric trimagnesic disilicate*	$\left\{\begin{array}{l}\mathbf{Si}\mathrm{HoMgo}''\\ \mathrm{Mgo}''\\ \mathbf{Si}\mathrm{HoMgo}''\end{array}\right..$
Steatite. *Trimagnesic tetrasilicate*	$\mathbf{Si}_4\mathrm{O}_5\mathrm{Mgo}''_3.$
Meerschaum. *Tetrahydric dimagnesic trisilicate*	$\left\{\begin{array}{l}\mathbf{Si}\mathrm{HoMgo}''\\ \mathrm{O}\\ \mathbf{Si}\mathrm{Ho}_2\\ \mathrm{O}\\ \mathbf{Si}\mathrm{HoMgo}''\end{array}\right..$
Pyrophyllite. *Dihydric aluminic tetrasilicate*	$\left.\begin{array}{l}\mathbf{Si}\mathrm{OHo}-\\ \mathbf{Si}\mathrm{O}-\\ \mathbf{Si}\mathrm{O}-\\ \mathbf{Si}\mathrm{OHo}-\end{array}\right]\mathrm{Alo}^{vi}.$
Anorthite. *Aluminic calcic disilicate*	$\mathbf{Si}_2({}'\mathbf{Al}'''_2\mathrm{O}_6)^{vi}\mathrm{Cao}''.$
Labradorite. *Aluminic calcic trisilicate*	$\left.\begin{array}{l}\mathbf{Si}\mathrm{O}-\\ \mathbf{Si}\mathrm{Cao}''-\\ \mathbf{Si}\mathrm{O}-\end{array}\right]\mathrm{Alo}^{vi}.$
Grossularia. *Aluminic tricalcic trisilicate*	$\left.\begin{array}{l}\mathbf{Si}\mathrm{Cao}''-\\ \mathbf{Si}\mathrm{Cao}''-\\ \mathbf{Si}\mathrm{Cao}''-\end{array}\right]\mathrm{Alo}^{vi}.$
Emerald. *Triglucinic aluminic hexasilicate*	$\mathbf{Si}_6\mathrm{O}_6\mathrm{Alo}^{vi}\mathrm{Glo}''_3.$
Chloropal. *Ferric trisilicate*	$\left.\begin{array}{l}\mathbf{Si}\mathrm{O}-\\ \mathbf{Si}\mathrm{O}-\\ \mathbf{Si}\mathrm{O}-\end{array}\right]\mathrm{Feo}^{vi}, 3\mathbf{O}\mathrm{H}_2,$
Felspar. Orthose. *Dipotassic aluminic hexasilicate*	$\mathbf{Si}_6\mathrm{O}_8\mathrm{Ko}_2\mathrm{Alo}^{vi}.$

SILICIC SULPHIDE.

$\mathbf{Si}\mathrm{S}''_2.$

Preparation.—By passing the vapour of carbonic disulphide over silicic anhydride heated to redness:—

$$\underset{\text{Silicic anhydride.}}{\mathbf{Si}\mathrm{O}_2} + \underset{\text{Carbonic disulphide.}}{\mathrm{CS}''_2} = \underset{\text{Silicic sulphide.}}{\mathbf{Si}\mathrm{S}''_2} + \underset{\text{Carbonic anhydride.}}{\mathrm{CO}_2}.$$

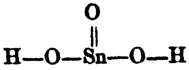
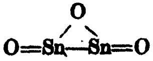
Reaction.—By the action of water, hydrosulphuric acid is evolved, and the solution contains silicic acid:—

$$\underset{\text{Silicic sulphide.}}{\mathbf{Si}S_2} + \underset{\text{Water.}}{4OH_2} = \underset{\text{Silicic acid.}}{\mathbf{Si}Ho_4} + \underset{\text{Sulphuretted hydrogen.}}{2SH_2.}$$

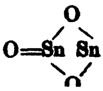
TIN, Sn.

Atomic weight =118. *Molecular weight unknown. Sp. gr.* 7·28. *Fuses at* 228°. *Atomicity* $''$ *and* ^{iv}, *also a pseudo-triad.*

The following are the names and probable formulæ of the principal compounds of this metal:—

Stannous chloride	$\mathbf{Sn}Cl_2$.	
Stannic chloride	$\mathbf{Sn}Cl_4$.	
Stannous oxide..............	$\mathbf{Sn}O$.	
Stannic oxide or anhydride	$\mathbf{Sn}O_2$.	
Distannous oxydichloride..	$\left\{\begin{array}{l}\mathbf{Sn}Cl\\O\\\mathbf{Sn}Cl.\end{array}\right.$	Cl—Sn—O—Sn—Cl
Stannous hydrate...........	$\mathbf{Sn}Ho_2$.	H—O—Sn—O—H
Stannic acid	$\mathbf{Sn}OHo_2$.	 $H—O—\overset{\overset{O}{\Vert}}{Sn}—O—H$
Dipotassic stannite	$\mathbf{Sn}Ko_2$.	
Dipotassic stannate.........	$\mathbf{Sn}OKo_2$, $4OH_2$.	
Distannic trioxide	$\left\{\begin{array}{l}\mathbf{Sn}O\\\mathbf{Sn}O\end{array}\right.O,$	

or

Stannous stannate	$\mathbf{Sn}OSno''$.	

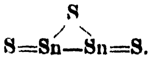
Metastannic acid (dried at 100°)......................
$$\left\{\begin{array}{l}\mathbf{Sn}Ho_3\\ O\\ \mathbf{Sn}Ho_2\\ O\\ \mathbf{Sn}O\\ O\\ \mathbf{Sn}Ho_2\\ O\\ \mathbf{Sn}Ho_3\end{array}\right.$$

Dipotassic metastannate...
$$\left\{\begin{array}{l}\mathbf{Sn}Ho_2Ko\\ O\\ \mathbf{Sn}Ho_2\\ O\\ \mathbf{Sn}O\\ O\\ \mathbf{Sn}Ho_2\\ O\\ \mathbf{Sn}Ho_2Ko\end{array}\right.$$

Stannous sulphide........... $\mathbf{Sn}S''$.

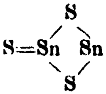
Stannic sulphide .... $\mathbf{Sn}S''_2$.

Distannic trisulphide ...... $\left\{\begin{array}{l}\mathbf{Sn}S''\\ \mathbf{Sn}S''\end{array}\right. S''$,

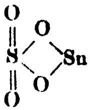


or

Stannous sulphostannate... $\mathbf{Sn}SSns''$.



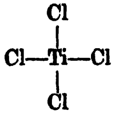
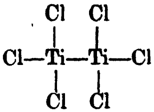
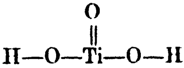
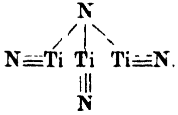
Stannous sulphate .... $\mathbf{S}O_2Sno''$.



TITANIUM, Ti.

Atomic weight =50. *Molecular weight unknown. Sp. gr.* 5·3. *Atomicity* '' *and* ^{iv}, *also a pseudo-triad.*

The following are the names and probable formulæ of the chief compounds of titanium :—

Titanic tetrachloride	$\mathbf{Ti}Cl_4$.	 $\begin{array}{c} Cl \\ \vert \\ Cl—Ti—Cl \\ \vert \\ Cl \end{array}$
Dititanic hexachloride ...	$\left\{\begin{array}{l}\mathbf{Ti}Cl_3\\ \mathbf{Ti}Cl_3\end{array}\right.$	 $\begin{array}{c} Cl \quad Cl \\ \vert \quad\; \vert \\ Cl—Ti—Ti—Cl \\ \vert \quad\; \vert \\ Cl \quad Cl \end{array}$
Titanous oxide..............	$\mathbf{Ti}O$.	$Ti{=}O$
Titanic oxide or anhydride (Rutile, Anatase, Brookite)	$\mathbf{Ti}O_2$.	
Titanic acid	$\mathbf{Ti}OHo_2$.	 $\begin{array}{c} O \\ \Vert \\ H—O—Ti—O—H \end{array}$
Titanic sulphide	$\mathbf{Ti}S''_2$.	
Dititanic dinitride	$\left\{\begin{array}{l}\mathbf{Ti}N'''\\ \mathbf{Ti}N'''\end{array}\right.$.	$N{\equiv}Ti—Ti{\equiv}N$
Trititanic tetranitride......	$\mathbf{Ti}_3N'''_4$	

CHAPTER XVI.

PENTAD ELEMENTS.

SECTION I. (*Continued from* Chapter XII.)

PHOSPHORUS, P_4.

Atomic weight =31. *Molecular weight* =124. *Molecular volume* ☐☐. 1 *litre of phosphorus vapour weighs* 62 *criths. Sp. gr.* 1·83. *Fuses at* 44–45°. *Boils at* 290°. *Atomicity* $'''$ *and* $^{\mathrm{v}}$. *Evidence of atomicity*:—

Phosphorous trihydride	$\mathbf{P}'''\mathrm{H}_3$.
Phosphorous trichloride	$\mathbf{P}'''\mathrm{Cl}_3$.
Phosphoric chloride	$\mathbf{P}^{\mathrm{v}}\mathrm{Cl}_5$.
Phosphonic iodide	$\mathbf{P}^{\mathrm{v}}\mathrm{H}_4\mathrm{I}$.

Occurrence.—In combination as a constituent of several minerals, and in small quantities in most rocks and soils.

In plants, and in the brain, nerves, urine, and bones of animals.

Manufacture.—Calcined bones or Sombrerite, both of which consist chiefly of calcic phosphate, are digested with sulphuric acid, by which the tricalcic diphosphate is converted into tetrahydric calcic diphosphate:—

$$\underset{\text{Tricalcic diphosphate (Bone-ash).}}{\mathbf{P}_2\mathrm{O}_2\mathrm{Cao}''_3} + \underset{\text{Sulphuric acid.}}{2\mathbf{S}\mathrm{O}_2\mathrm{Ho}_2} = \underset{\text{Tetrahydric calcic diphosphate.}}{\mathbf{P}_2\mathrm{O}_2\mathrm{Ho}_4\mathrm{Cao}''} + \underset{\text{Calcic sulphate.}}{2\mathbf{S}\mathrm{O}_2\mathrm{Cao}''}.$$

The tetrahydric calcic phosphate is extracted with water from the calcic sulphate, evaporated, mixed with charcoal, dried and distilled, when phosphorus, carbonic oxide, and tricalcic diphosphate are produced:—

$$\underset{\text{Calcic metaphosphate.}}{3\mathbf{P}_2\mathrm{O}_4\mathrm{Cao}''} + \mathrm{C}_{10} = \underset{\text{Tricalcic diphosphate.}}{\mathbf{P}_2\mathrm{O}_2\mathrm{Cao}''_3} + \underset{\text{Carbonic oxide.}}{10\mathbf{C}\mathrm{O}} + \mathrm{P}_4.$$

AMORPHOUS PHOSPHORUS. *Allotropic Phosphorus. Red Phosphorus.*

Obtained by heating common phosphorus to 230°–250° in close vessels.

Neither the number nor the arrangement of the atoms in the molecule of this variety of phosphorus is known.

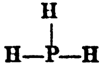
COMPOUNDS OF PHOSPHORUS WITH HYDROGEN.

Phosphorus forms three compounds with hydrogen. They cannot be obtained by the direct combination of their elements.

Solid phosphoretted hydrogen ...	$\left\{\begin{array}{l}\mathbf{P}(P'''H)''\\ \mathbf{P}(P'''H)''\end{array}\right.$?
Liquid ditto	$'\mathbf{P}''_2H_4$.
Gaseous ditto	$\mathbf{P}H_3$.

GASEOUS PHOSPHORETTED HYDROGEN.

$\mathbf{P}H_3$



$$\begin{array}{c} H \\ | \\ H—P—H \end{array}$$

Molecular weight =34. *Molecular volume* ☐☐. 1 *litre weighs* 17 *criths.*

Preparation.—1. By heating hypophosphorous acid:—

$$\underset{\text{Hypophosphorous acid.}}{2\mathbf{P}OH_2Ho} = \underset{\text{Phosphoretted hydrogen.}}{\mathbf{P}H_3} + \underset{\text{Phosphoric acid.}}{\mathbf{P}OHo_3}.$$

2. By heating phosphorous acid:—

$$\underset{\text{Phosphorous acid.}}{4\mathbf{P}OHHo_2} = \underset{\text{Phosphoretted hydrogen.}}{\mathbf{P}H_3} + \underset{\text{Phosphoric acid.}}{3\mathbf{P}OHo_3}.$$

3. By heating phosphorus with solution of sodic or potassic drate :—

$$\underset{\substack{\text{Sodic}\\ \text{hydrate.}}}{3\mathbf{O}NaH} + P_4 + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\substack{\text{Sodic}\\ \text{hypophosphite.}}}{3\mathbf{P}OH_2Nao} + \underset{\substack{\text{Phosphoretted}\\ \text{hydrogen.}}}{\mathbf{P}H_3}.$$

The gas prepared by this process contains free hydrogen and the vapour of liquid phosphoretted hydrogen.

Reactions.—1. By combustion in oxygen it yields phosphoric acid :—

$$\underset{\substack{\text{Phosphoretted}\\ \text{hydrogen.}}}{\mathbf{P}H_3} + 2O_2 = \underset{\substack{\text{Phosphoric}\\ \text{acid.}}}{\mathbf{P}OHo_3}.$$

2. When passed through a solution of cupric sulphate, it causes a black precipitate of cupric phosphide :—

$$\underset{\substack{\text{Phosphoretted}\\ \text{hydrogen.}}}{2\mathbf{P}H_3} + \underset{\substack{\text{Cupric}\\ \text{sulphate.}}}{3\mathbf{S}O_2Cuo''} = \underset{\substack{\text{Cupric}\\ \text{phosphide.}}}{\mathbf{P}_2Cu''_3}, \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{3\mathbf{S}O_2Ho_2}.$$

3. When passed through a solution of argentic nitrate, metallic silver and nitric and phosphoric acids are formed :—

$$\underset{\substack{\text{Phosphoretted}\\ \text{hydrogen.}}}{\mathbf{P}H_3} + \underset{\substack{\text{Argentic}\\ \text{nitrate.}}}{8\mathbf{N}O_2Ago} + \underset{\text{Water.}}{4\mathbf{O}H_2} = \underset{\substack{\text{Phosphoric}\\ \text{acid.}}}{\mathbf{P}OHo_3} + 4Ag_2 + \underset{\substack{\text{Nitric}\\ \text{acid.}}}{8\mathbf{N}O_2Ho}.$$

4. It unites directly with hydriodic and hydrobromic acids when they are presented to it in the nascent state, forming compounds isomorphous with the corresponding substances in the nitrogen series :—

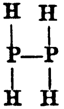
$$\underset{\substack{\text{Phosphoretted}\\ \text{hydrogen.}}}{\mathbf{P}H_3} + 3I_2 = \underset{\substack{\text{Phosphorous}\\ \text{triiodide.}}}{\mathbf{P}I_3} + \underset{\substack{\text{Hydriodic}\\ \text{acid.}}}{3HI};$$

$$\underset{\substack{\text{Phosphoretted}\\ \text{hydrogen.}}}{3\mathbf{P}H_3} + \underset{\substack{\text{Hydriodic}\\ \text{acid.}}}{3HI} = \underset{\substack{\text{Phosphonic}\\ \text{iodide.}}}{3\mathbf{P}H_4I}.$$

In this behaviour phosphoretted hydrogen bears a striking analogy to ammonia, although, unlike the latter compound, it does not unite with other acids.

LIQUID PHOSPHORETTED HYDROGEN.

$'\mathbf{P}''_2H_4$, or $\left\{\begin{matrix}\mathbf{P}H_2\\ \mathbf{P}H_2\end{matrix}\right.$.



Molecular weight = 66. *Boils between* 30° *and* 40°.

Preparation.—By the action of water or very dilute hydrochloric acid upon calcic phosphide, $'\mathbf{P}''_2Ca''_2$, the gas evolved being transmitted through a freezing-mixture:—

$$\underset{\text{Calcic phosphide.}}{'\mathbf{P}''_2Ca''_2} + \underset{\text{Water.}}{4OH_2} = \underset{\text{Liquid phosphoretted hydrogen.}}{'\mathbf{P}''_2H_4} + \underset{\text{Calcic hydrate.}}{2\mathbf{Ca}''Ho_2}.$$

The calcic phosphide is prepared by passing the vapour of phosphorus over lime heated to redness:—

$$14P + \underset{\text{Lime.}}{14\mathbf{Ca}O} = \underset{\text{Calcic pyrophosphate.}}{2\mathbf{P}_2O_3Cao''_2} + \underset{\text{Calcic phosphide.}}{5'\mathbf{P}''_2Ca''_2}.$$

Reaction.—Decomposed by sunlight into solid and gaseous phosphoretted hydrogen:—

$$\underset{\text{Liquid phosphoretted hydrogen.}}{5'\mathbf{P}''_2H_4} = \underset{\text{Gaseous phosphoretted hydrogen.}}{6\mathbf{P}H_3} + \underset{\text{Solid phosphoretted hydrogen.}}{\left\{\begin{matrix}\mathbf{P}(P'''H)''\\ \mathbf{P}(P'''H)''\end{matrix}\right. ?}$$

SOLID PHOSPHORETTED HYDROGEN.

$$\left\{\begin{matrix}\mathbf{P}(P'''H)''\\ \mathbf{P}(P'''H)''\end{matrix}\right. ?$$

Molecular weight = 126 ?

Preparation.—By dissolving calcic phosphide in concentrated hydrochloric acid, or by the action of light upon the liquid phosphoretted hydrogen.

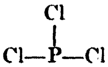
COMPOUNDS OF PHOSPHORUS WITH CHLORINE.

Phosphorus forms two compounds with chlorine :—

Phosphorous trichloride	$\mathbf{P}Cl_3$.
Phosphoric chloride	$\mathbf{P}Cl_5$.

PHOSPHOROUS CHLORIDE.

$\mathbf{P}Cl_3$.



$$\begin{array}{c} Cl \\ | \\ Cl—P—Cl \end{array}$$

Molecular weight = 137·5. *Molecular volume* . 1 *litre of phosphorous trichloride vapour weighs* 68·75 *criths. Sp. gr.* 1·45. *Boils at* 74°.

Preparation.—By the action of chlorine upon phosphorus :—

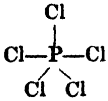
$$P_2 \quad + \quad 3Cl_2 \quad = \quad 2\mathbf{P}Cl_3.$$

Reaction.—By the action of water it yields hydrochloric and phosphorous acids :—

$$\underset{\text{Phosphorous trichloride.}}{\mathbf{P}Cl_3} \quad + \quad \underset{\text{Water.}}{3\mathbf{O}H_2} \quad = \quad \underset{\text{Hydrochloric acid.}}{3HCl} \quad + \quad \underset{\text{Phosphorous acid.}}{\mathbf{P}OHHo_2}.$$

PHOSPHORIC CHLORIDE.

$\mathbf{P}Cl_5$.



Molecular weight = 208·5. *Molecular volume* ☐☐ *to* ⊞.

1 *litre of phosphoric chloride weighs* 52·1 *to* 104·25 *criths. Volatilizes below* 100°.

Preparation.—By the action of chlorine upon phosphorous trichloride :—

$$\underset{\substack{\text{Phosphorous}\\\text{trichloride.}}}{\mathbf{P}Cl_3} + Cl_2 = \underset{\substack{\text{Phosphoric}\\\text{chloride.}}}{\mathbf{P}Cl_5}.$$

Reactions.—1. By the action of an excess of water it produces hydrochloric acid and phosphoric acid :—

$$\underset{\substack{\text{Phosphoric}\\\text{chloride.}}}{\mathbf{P}Cl_5} + \underset{\text{Water.}}{4\mathbf{O}H_2} = \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{5HCl} + \underset{\substack{\text{Phosphoric}\\\text{acid.}}}{\mathbf{PO}Ho_3}.$$

2. When submitted to the action of alcohols and acids, the chlorides of the radicals of the alcohols and acids are obtained, thus :—

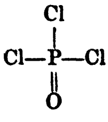
$$\underset{\substack{\text{Ethylic}\\\text{alcohol.}}}{\left\{\begin{matrix}\mathbf{C}H_3\\\mathbf{C}H_2Ho\end{matrix}\right.} + \underset{\substack{\text{Phosphoric}\\\text{chloride.}}}{\mathbf{P}Cl_5} = \underset{\substack{\text{Ethylic}\\\text{chloride.}}}{\left\{\begin{matrix}\mathbf{C}H_3\\\mathbf{C}H_2Cl\end{matrix}\right.} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{HCl} + \underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{\mathbf{PO}Cl_3}.$$

$$\underset{\text{Acetic acid.}}{\left\{\begin{matrix}\mathbf{C}H_3\\\mathbf{CO}Ho\end{matrix}\right.} + \underset{\substack{\text{Phosphoric}\\\text{chloride.}}}{\mathbf{P}Cl_5} = \underset{\substack{\text{Acetylic}\\\text{chloride.}}}{\left\{\begin{matrix}\mathbf{C}H_3\\\mathbf{CO}Cl\end{matrix}\right.} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{HCl} + \underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{\mathbf{PO}Cl_3}.$$

COMPOUND OF PHOSPHORUS WITH CHLORINE AND OXYGEN.

PHOSPHORIC OXYTRICHLORIDE.

$\mathbf{PO}Cl_3$.



$$\begin{matrix} & Cl & \\ & | & \\ Cl— & P & —Cl \\ & \| & \\ & O & \end{matrix}$$

Molecular weight **=153·5.** ***Molecular volume*** ☐☐. **1** ***litre of***

phosphoric oxytrichloride vapour weighs 76·75 *criths. Sp. gr.* 1·7. *Boiling-point* 110°.

Preparation.—1. By the action of a limited quantity of water on phosphoric chloride:—

$$\underset{\substack{\text{Phosphoric}\\\text{chloride.}}}{\mathbf{P}Cl_5} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{\mathbf{P}OCl_3} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{2HCl.}$$

2. By passing oxygen through boiling phosphorous trichloride:—

$$\underset{\substack{\text{Phosphorous}\\\text{trichloride.}}}{\mathbf{P}Cl_3} + O = \underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{\mathbf{P}OCl_3.}$$

3. By heating phosphoric chloride with phosphoric anhydride:—

$$\underset{\substack{\text{Phosphoric}\\\text{anhydride.}}}{\mathbf{P}_2O_5} + \underset{\substack{\text{Phosphoric}\\\text{chloride.}}}{3\mathbf{P}Cl_5} = \underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{5\mathbf{P}OCl_3.}$$

4. It is formed as a secondary product in the preparation of the chlorides of alcohol and acid radicals as above described (p. 116).

Reactions.—1. By contact with water it is transformed into hydrochloric and phosphoric acids:—

$$\underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{\mathbf{P}OCl_3} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\substack{\text{Phosphoric}\\\text{acid.}}}{\mathbf{P}OHo_3} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{3HCl.}$$

2. By distillation with the salts of organic acids it yields the chloracids:—

$$\underset{\substack{\text{Sodic}\\\text{acetate.}}}{3\left\{\begin{matrix}\mathbf{C}H_3\\\mathbf{C}ONao\end{matrix}\right.} + \underset{\substack{\text{Phosphoric}\\\text{oxytrichloride.}}}{\mathbf{P}OCl_3} = \underset{\substack{\text{Acetylic}\\\text{chloride.}}}{3\left\{\begin{matrix}\mathbf{C}H_3\\\mathbf{C}OCl\end{matrix}\right.} + \underset{\substack{\text{Sodic}\\\text{phosphate.}}}{\mathbf{P}ONao_3.}$$

COMPOUND OF PHOSPHORUS WITH CHLORINE AND SULPHUR.

PHOSPHORIC SULPHOTRICHLORIDE.

$$\mathbf{P}S''Cl_3.$$

Molecular weight =169·5. *Boils at* 128°.

Preparation.—By the action of sulphuretted hydrogen upon phosphoric chloride:—

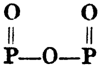
$$\underset{\text{Phosphoric chloride.}}{\mathbf{P}Cl_5} + \underset{\text{Sulphuretted hydrogen.}}{\mathbf{S}H_2} = \underset{\text{Phosphoric sulphotrichloride.}}{\mathbf{P}S''Cl_3} + \underset{\text{Hydrochloric acid.}}{2HCl.}$$

Reaction.—When boiled with sodic hydrate, it yields sodic chloride and trisodic sulphophosphate:—

$$\underset{\text{Sodic hydrate.}}{6\mathbf{O}NaH} + \underset{\text{Phosphoric sulphotrichloride.}}{\mathbf{P}S''Cl_3} = \underset{\text{Sodic chloride.}}{3NaCl} + \underset{\text{Trisodic sulpho-phosphate.}}{\mathbf{P}S''Nao_3} + \underset{\text{Water.}}{3\mathbf{O}H_2.}$$

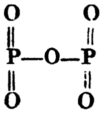
COMPOUNDS OF PHOSPHORUS WITH OXYGEN AND HYDROXYL.

Phosphorous anhydride ... $\mathbf{P}_2O_3$.



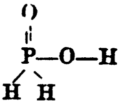
$$\begin{array}{ccc} O & & O \\ \| & & \| \\ P & —O— & P \end{array}$$

Phosphoric anhydride...... $\mathbf{P}_2O_5$.

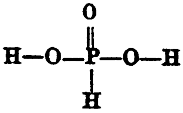


$$\begin{array}{ccc} O & & O \\ \| & & \| \\ P & —O— & P \\ \| & & \| \\ O & & O \end{array}$$

Hypophosphorous acid ... $\mathbf{P}OH_2Ho$.

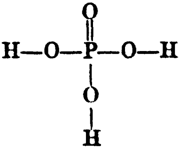


Phosphorous acid............ $\mathbf{P}OHHo_2$.



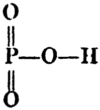
$$\begin{array}{ccccc} & & O & & \\ & & \| & & \\ H & —O— & P & —O— & H \\ & & | & & \\ & & H & & \end{array}$$

Phosphoric acid (tribasic) ... } $\mathbf{P}OHo_3$.



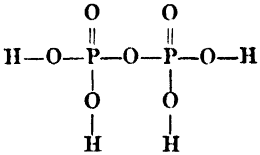
$$\mathrm{H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{H}{|}}{O}}{\underset{|}{P}}}{-}O{-}H}$$

Metaphosphoric acid (monobasic). } $\mathbf{P}O_2Ho$.



$$\mathrm{\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{P}}{-}O{-}H}$$

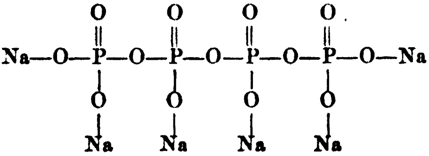
Pyrophosphoric acid (tetra-basic) .... } $\mathbf{P}_2O_3Ho_4$.



$$\mathrm{H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{H}{|}}{O}}{\underset{|}{P}}}{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{H}{|}}{O}}{\underset{|}{P}}}{-}O{-}H}$$

Hexabasic phosphoric acid .... $\mathbf{P}_4O_7Ho_6$.

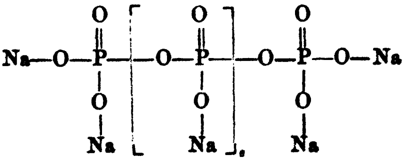
Sodium salt (Fleitmann and Henneberg) (*Hexasodic tetraphosphate*) .... $\mathbf{P}_4O_7Nao_6$.



$$\mathrm{Na{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}O{-}Na}$$

Dodecabasic phosphoric acid .... $\mathbf{P}_{10}O_{19}Ho_{12}$.

Sodium salt (Fleitmann and Henneberg) (*Dodecasodic decaphosphate*) .... $\mathbf{P}_{10}O_{19}Nao_{12}$.



$$\mathrm{Na{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}\left[{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}\right]_8{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{Na}{|}}{O}}{\underset{|}{P}}}{-}O{-}Na}$$

PHOSPHOROUS ANHYDRIDE.

P_2O_3.

Molecular weight =110.

Preparation.—By the slow oxidation of phosphorus in a gentle current of dry air.

Reaction.—In contact with water it produces phosphorous acid:—

$$\underset{\text{Phosphorous anhydride.}}{P_2O_3} + \underset{\text{Water.}}{3OH_2} = \underset{\text{Phosphorous acid.}}{2POHHo_2}.$$

PHOSPHOROUS ACID.

$POHHo_2$.

Molecular weight =82.

Preparation.—1. By the action of water on phosphorous anhydride as above.

2. By the slow oxidation of phosphorus in moist air.

3. By the action of water upon phosphorous chloride (see p. 115).

4. By passing chlorine through phosphorus under hot water.

Reactions.—1. When heated, it yields phosphoric acid and phosphoretted hydrogen:—

$$\underset{\text{Phosphorous acid.}}{4POHHo_2} = \underset{\text{Phosphoric acid.}}{3POHo_3} + \underset{\text{Phosphoretted hydrogen.}}{PH_3}$$

2. It absorbs oxygen from the air, yielding phosphoric acid:—

$$\underset{\text{Phosphorous acid.}}{2POHHo_2} + O_2 = \underset{\text{Phosphoric acid.}}{2POHo_3}.$$

PHOSPHORIC ANHYDRIDE.

P_2O_5.

Molecular weight =142.

Preparation.—By burning phosphorus in excess of dry air or oxygen.

Reaction.—By contact with water it forms metaphosphoric acid:—

$$\underset{\text{Phosphoric anhydride.}}{P_2O_5} + \underset{\text{Water.}}{OH_2} = \underset{\text{Metaphosphoric acid.}}{2PO_2Ho.}$$

METAPHOSPHORIC ACID.

PO_2Ho.

Molecular weight =80.

Preparation.—1. By dissolving phosphoric anhydride in water (see above).

2. By heating phosphoric acid to redness:—

$$\underset{\text{Phosphoric acid.}}{POHo_3} = \underset{\text{Metaphosphoric acid.}}{PO_2Ho} + \underset{\text{Water.}}{OH_2.}$$

Preparation of metaphosphates.—The metaphosphates may be produced—

1. By igniting a dihydric phosphate with a fixed base:—

$$\underset{\text{Dihydric sodic phosphate.}}{POHo_2Nao} = \underset{\text{Sodic metaphosphate.}}{PO_2Nao} + \underset{\text{Water.}}{OH_2.}$$

2. By igniting a monohydric phosphate which contains one atom of a volatile base:—

$$\underset{\text{Hydric sodic ammonic phosphate.}}{POHoNao(N^{v}H_4O)} = \underset{\text{Sodic metaphosphate.}}{PO_2Nao} + \underset{\text{Ammonia.}}{NH_3} + \underset{\text{Water.}}{OH_2.}$$

3. By igniting a dihydric pyrophosphate:—

$$\underset{\text{Dihydric disodic pyrophosphate.}}{P_2O_3Ho_2Nao_2} = \underset{\text{Sodic metaphosphate.}}{2PO_2Nao} + \underset{\text{Water.}}{OH_2.}$$

PYROPHOSPHORIC ACID.

$P_2O_3Ho_4$.

Molecular weight =178.

Preparation.—1. By decomposing plumbic pyrophosphate by hydrosulphuric acid:—

$$\underset{\text{Plumbic}}{P_2O_3Pbo''_2} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{2SH_2} = \underset{\substack{\text{Plumbic}\\\text{sulphide.}}}{2PbS''} + \underset{\substack{\text{Pyrophosphoric}\\\text{acid.}}}{P_2O_3Ho_4}.$$

2. By heating tribasic phosphoric acid to 213°.

Pyrophosphates are prepared by heating monohydric phosphates containing two atoms of a monad metal:—

$$\underset{\substack{\text{Hydric disodic}\\\text{phosphate.}}}{2POHoNao_2} = \underset{\substack{\text{Sodic}\\\text{pyrophosphate.}}}{P_2O_3Nao_4} + \underset{\text{Water.}}{OH_2}.$$

PHOSPHORIC ACID, *Tribasic Phosphoric Acid.*

$POHo_3$.

Molecular weight =98.

Preparation.—1. By boiling a solution of phosphoric anhydride or of metaphosphoric acid in water:—

$$\underset{\substack{\text{Phosphoric}\\\text{anhydride.}}}{P_2O_5} + \underset{\text{Water.}}{3OH_2} = \underset{\substack{\text{Phosphoric}\\\text{acid.}}}{2POHo_3}.$$

2. By the oxidation of amorphous phosphorus with nitric acid, and then boiling the product with water.

3. By the action of water upon phosphoric chloride and phosphoric oxytrichloride (see pp. 116 and 117).

4. By the combustion of phosphoretted hydrogen in air or oxygen:—

$$PH_3 + 2O_2 = POHo_3.$$

5. By decomposing tricalcic diphosphate (bone-ash) with a large excess of sulphuric acid:—

$$\underset{\substack{\text{Tricalcic}\\\text{diphosphate.}}}{\mathbf{P}_2O_2Cao''_3} + \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{3\mathbf{S}O_2Ho_2} + \underset{\text{Water.}}{6\mathbf{O}H_2} = \underset{\substack{\text{Phosphoric}\\\text{acid.}}}{2\mathbf{P}OHo_3}$$

$$+ \underset{\substack{\text{Gypsum. Tetrahydric}\\\text{calcic sulphate.}}}{3\mathbf{S}Ho_4Cao''}.$$

Reaction.—When heated to 213° it produces pyrophosphoric acid:—

$$\underset{\substack{\text{Phosphoric}\\\text{acid.}}}{2\mathbf{P}OHo_3} = \underset{\substack{\text{Pyrophosphoric}\\\text{acid.}}}{\mathbf{P}_2O_3Ho_4} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

The phosphates are a numerous and important class of salts. The following list contains some of the most interesting:—

Common sodic phosphate (*Hydric disodic phosphate*)	$\mathbf{P}OHoNao_2$, $12\mathbf{O}H_2$.
Trisodic phosphate	$\mathbf{P}ONao_3$, $12\mathbf{O}H_2$.
Hydric sodic potassic phosphate	$\mathbf{P}OHoNaoKo$, $\mathbf{O}H_2$.
Apatite (*Francolite*)	$\mathbf{P}_3O_3Cao''_4\left(\begin{smallmatrix}O\\F\end{smallmatrix}Ca''\right)$.
Triple phosphate (*Ammonic magnesic phosphate*)	$\mathbf{P}OAmoMgo''$,
Vivianite	$\mathbf{P}_2O_2Feo''_3$, $8\mathbf{O}H_2$.
Wavellite	$\mathbf{P}_4O('Al'''_2O_6)^{vi}_3$, $12\mathbf{O}H_2$.
Pyromorphite	$\mathbf{P}_3O_3Pb$

VANADIUM, V_4?

Atomic weight =51·3. *Probable molecular weight* =205·2. *Atomicity* ''' and v. *Evidence of atomicity*:—

Vanadous chloride	$\mathbf{V}'''Cl_3$.
Vanadic oxychloride	$\mathbf{V}^vOCl_3$.

Occurrence.—Sparingly in some lead- and iron-ores.

Preparation.—By heating to bright redness vanadous chloride in a current of dry hydrogen.

$$\underset{\text{Vanadous chloride.}}{2\mathbf{V}Cl_3} + 3H_2 = \mathbf{V}_2 + \underset{\text{Hydrochloric acid.}}{6HCl.}$$

The following list contains the names and formulæ of the remaining chief compounds of vanadium:—

Hypovanadous chloride	$\left\{\begin{matrix}\mathbf{V}Cl_2\\ \mathbf{V}Cl_2\end{matrix}\right.$.
Hypovanadous oxide	$\mathbf{V}_2O_2$.
Hypovanadic chloride	$\mathbf{V}Cl_4$ or $\left\{\begin{matrix}\mathbf{V}Cl_4\\ \mathbf{V}Cl_4\end{matrix}\right.$.
Vanadous oxide	
Hypovanadic oxide	
Vanadic anhydride..............	$\mathbf{V}_2O_5$.
Metavanadic acid	$\mathbf{V}O_2Ho$.
Tribasic vanadic acid	$\mathbf{V}OHo_3$.
Pyrovanadic acid	$\mathbf{V}_2O_3Ho_4$.

ARSENIC, As_4.

Atomic weight =75. *Molecular weight* =300. *Molecular volume* □□. 1 *litre of arsenic vapour weighs* 150 *criths.* *Sp. gr.* 5·6 to 5·9. *Volatile at* 180°. *Atomicity* $'''$ and ^{v}. *Evidence of atomicity*:—

Arseniuretted hydrogen	$\mathbf{As}'''H_3$.
Arsenious chloride	$\mathbf{As}'''Cl_3$.
Tetrethylarsonic chloride	$\mathbf{As}^{v}Et_4Cl$.

Occurrence.—Abundantly in nature, in various ores, and sometimes in the free state. In some mineral waters, and in the water and mud of many rivers. In coal-smoke, and consequently in the air of towns.

Preparation.—By reducing, with charcoal, arsenious anhydride, which is produced in the roasting of many ores:—

$$\underset{\substack{\text{Arsenious}\\ \text{anhydride.}}}{\mathbf{As}_2O_3} + 3C = \mathbf{As}_2 + \underset{\substack{\text{Carbonic}\\ \text{oxide.}}}{3CO.}$$

COMPOUND OF ARSENIC WITH HYDROGEN.

ARSENIURETTED HYDROGEN, *Arsenious Hydride.*

$\mathbf{As}H_3$.

Molecular weight = 78. *Molecular volume* ☐☐. 1 *litre weighs* 39 *criths. Boils at* −40°.

Preparation.—1. In the pure state by the action of sulphuric acid on an alloy of arsenic and zinc:—

$$\underset{\substack{\text{Arsenious}\\\text{zincoide.}}}{\mathbf{As}_2Zn''_3} + \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{3\mathbf{S}O_2Ho_2} = \underset{\substack{\text{Zincic}\\\text{sulphate.}}}{3\mathbf{S}O_2Zno''} + \underset{\substack{\text{Arseniuretted}\\\text{hydrogen.}}}{2\mathbf{As}H_3}.$$

2. By the action of nascent hydrogen upon soluble arsenic compounds, as by the introduction of arsenious acid into an apparatus evolving hydrogen:—

$$\underset{\substack{\text{Arsenious}\\\text{acid.}}}{\mathbf{As}Ho_3} + 3H_2 = \underset{\substack{\text{Arseniuretted}\\\text{hydrogen.}}}{\mathbf{As}H_3} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

Reactions.—1. When burnt with free access of air, it gives water and arsenious anhydride:—

$$\underset{\substack{\text{Arseniuretted}\\\text{hydrogen.}}}{2\mathbf{As}H_3} + 3O_2 = \underset{\substack{\text{Arsenious}\\\text{anhydride.}}}{\mathbf{As}_2O_3} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

2. When burnt with a limited supply of air, it yields water and free arsenic:—

$$\underset{\substack{\text{Arseniuretted}\\\text{hydrogen.}}}{} + 3O_2 = As_4 + \underset{\text{Water.}}{6\mathbf{O}H_2}.$$

3. When exposed to a red heat, it is decomposed into arsenic and hydrogen.

4. Passed through a solution of argentic nitrate, it yields

a precipitate of metallic silver, arsenious and nitric acids remaining in solution :—

$$\underset{\text{Argentic nitrate.}}{6\mathbf{N}O_2Ago} + \underset{\text{Water.}}{3\mathbf{O}H_2} + \underset{\text{Arseniuretted hydrogen.}}{\mathbf{As}H_3} = \underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho}$$

$$+ \underset{\text{Arsenious acid.}}{\mathbf{As}Ho_3} + 3Ag_2.$$

COMPOUND OF ARSENIC WITH CHLORINE.

ARSENIOUS CHLORIDE.

Molecular weight =181·5. *Molecular volume* ☐☐. 1 *litre of arsenious chloride vapour weighs* 90·75 *criths. Sp gr.* 2·205. *Boils at* 132°.

Preparation.—1. By the action of dry chlorine upon arsenic :—

$$As_2 + 3Cl_2 = \underset{\text{Arsenious chloride.}}{2\mathbf{As}Cl_3}$$

2. By distilling arsenic with mercuric chloride (corrosive sublimate) :—

$$As_2 + \underset{\text{Mercuric chloride.}}{6\mathbf{Hg}Cl_2} = \underset{\text{Mercurous chloride.}}{3'\mathbf{Hg}'_2Cl_2} + \underset{\text{Arsenious chloride.}}{2\mathbf{As}Cl_3}$$

3. By distilling sodic chloride, arsenious anhydride, and sulphuric acid :—

$$\underset{\text{Arsenious anhydride.}}{\mathbf{As}_2O_3} + \underset{\text{Sodic chloride.}}{6NaCl} + \underset{\text{Sulphuric acid.}}{6\mathbf{S}O_2Ho_2} = \underset{\text{Arsenious chloride.}}{2\mathbf{As}Cl_3}$$

$$+ \underset{\text{Hydric sodic}}{6\mathbf{S}O_2HoNao} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

Reaction.—With excess of water it forms arsenious and hydrochloric acids:—

$$\underset{\text{Arsenious chloride.}}{\mathbf{As}Cl_3} + \underset{\text{Water.}}{3OH_2} = \underset{\text{Hydrochloric acid.}}{3HCl} + \underset{\text{Arsenious acid.}}{\mathbf{As}Ho_3}.$$

COMPOUNDS OF ARSENIC WITH OXYGEN AND HYDROXYL.

Arsenious anhydride	$(\mathbf{As}_2O_3)_2$.
Arsenic anhydride	$\mathbf{As}_2O_5$.
Arsenious acid	$\mathbf{As}Ho_3$.
Arsenic acid	$\mathbf{As}OHo_3$.

ARSENIOUS ANHYDRIDE, *Arsenic, White Arsenic, White Oxide of Arsenic.*

$(\mathbf{As}_2O_3)_2$.

Molecular weight = 396. *Molecular volume* □□. 1 *litre of arsenious anhydride vapour weighs* 198 *criths.* *Sp. gr.* 3·7.

Occurrence.—Very rare in nature.

Preparation.—1. By burning arsenic in air or oxygen.

2. By roasting arsenical ores in certain metallurgical operations.

ARSENIOUS ACID.

$\mathbf{As}Ho_3$.

Molecular weight = 126.

Only known in solution.

Arsenious acid forms many unstable and difficultly crystallizable salts, of which the following are examples:—

Dihydric potassic arsenite (*Fowler's solution*)	$\mathbf{As}Ho_2Ko$.
Hydric cupric arsenite (*Scheele's green*)	$\mathbf{As}HoCuo''$.
Triargentic arsenite	$\mathbf{As}Ago_3$.

A monobasic arsenious acid, $\mathbf{As}OHo$, corresponding to nitrous acid, appears to exist, one of its compounds, $\mathbf{As}OAmo$, being known: Arsenious acid when boiled with cupric acetate yields Schweinfurt green, $3\mathbf{As}_2O_2Cuo''$, $\mathbf{Cu}(C_2H_3O_2)_2$.

ARSENIC ANHYDRIDE.

$\mathbf{As}_2O_5$,

Molecular weight =230.

Preparation.—By heating arsenic acid nearly to redness:—

$2\mathbf{As}OHo_3$	$=$	$3\mathbf{O}H_2$	$+$	$\mathbf{As}_2O_5$.
Arsenic acid.		Water.		Arsenic anhydride.

ARSENIC ACID.

$\mathbf{As}OHo_3$.

Molecular weight =142.

Preparation.—By treating arsenious anhydride with nitric acid:—

$\mathbf{As}_2O_3$	$+$	$2\mathbf{N}O_2Ho$	$+$	$2\mathbf{O}H_2$	$=$	$2\mathbf{As}OHo_3$	$+$	$\mathbf{N}_2O_3$.
Arsenious anhydride.		Nitric acid.		Water.		Arsenic acid.		Nitrous anhydride.

Salts are known which are derived from three acids of the following formulæ:—

$\mathbf{As}O_2Ho$,	$\mathbf{As}OHo_3$,	$\mathbf{As}_2O_3Ho_4$
Metarsenic acid.	Arsenic acid.	Pyrarsenic acid.
corresponding to	corresponding to	corresponding to
$\mathbf{P}O_2Ho$,	$\mathbf{P}OHo_3$,	$\mathbf{P}_2O_3Ho_4$,
Metaphosphoric acid.	Phosphoric acid.	Pyrophosphoric acid.
and		
$\mathbf{N}O_2Ho$.		
Nitric acid.		

COMPOUNDS OF ARSENIC WITH SULPHUR AND HYDROSULPHYL.

Realgar	$\left\{\begin{matrix}\mathbf{As}S'' \\ \mathbf{As}S''\end{matrix}\right. = {}'\mathbf{As}''_2S''_2.$
Sulpharsenious anhydride (*Arsenious sulphide*)	$\mathbf{As}_2S''_3.$
Sulpharsenic anhydride (*Arsenic sulphide*)	$\mathbf{As}_2S''_5.$
Sulpharsenious acid	$\mathbf{As}Hs_3.$
Sulpharsenic acid	$\mathbf{As}S''Hs_3.$

REALGAR, *Diarsenious Disulphide.*

$$ {}'\mathbf{As}''_2S''_2, \text{ or } \left\{\begin{matrix}\mathbf{As}S'' \\ \mathbf{As}S''\end{matrix}\right. .$$

Molecular weight =214. *Sp. gr.* 3·5.

Occurrence.—Found native.

Preparation.—By heating sulphur with arsenious anhydride:—

$$S_7 + \underset{\text{Arsenious anhydride.}}{2\mathbf{As}_2O_3} = \underset{\text{Sulphurous anhydride.}}{3\mathbf{S}O_2} + \underset{\text{Realgar.}}{2'\mathbf{As}''_2\mathbf{S}''_2}.$$

SULPHARSENIOUS ANHYDRIDE, *Arsenious Sulphide, Orpiment.*

$$\mathbf{As}_2S''_3.$$

Molecular weight =246. *Sp. gr.* 3·5.

Occurrence.—Found native.

Preparation.—By passing sulphuretted hydrogen through a solution of arsenious anhydride in hydrochloric acid:—

$$\underset{\text{Arsenious chloride.}}{2\mathbf{As}Cl_3} + \underset{\text{Sulphuretted hydrogen.}}{3\mathbf{S}H_2} = \underset{\text{Hydrochloric acid.}}{6HCl} + \underset{\text{Arsenious sulphide.}}{\mathbf{As}_2S''_3}.$$

Reaction.—Arsenious sulphide dissolves in caustic alkali, producing an arsenite and a sulpharsenite:—

$$\underset{\text{Arsenious sulphide.}}{As_2S''_3} + \underset{\text{Potassic hydrate.}}{4OKH} = \underset{\text{Hydric dipotassic arsenite.}}{AsHoKo_2} + \underset{\text{Sulphhydric disulphopotassic sulpharsenite.}}{AsHsKs_2} + \underset{\text{Water.}}{OH_2}$$

By the addition of an acid, the arsenious sulphide is reprecipitated:—

$$\underset{\text{Hydric dipotassic arsenite.}}{AsHoKo_2} + \underset{\text{Sulphhydric disulphopotassic sulpharsenite.}}{AsHsKs_2} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Potassic chloride.}}{4KCl}$$

$$+ \underset{\text{Arsenious sulphide.}}{As_2S''_3} + \underset{\text{Water.}}{3OH_2}.$$

Proustite is a native trisulphargentic sulpharsenite, $AsAgs_3$.

SULPHARSENIC ANHYDRIDE, *Arsenic Sulphide.*

$As_2S''_5$.

Molecular weight =310.

Preparation.—By fusing together arsenious sulphide and sulphur. Sulpharsenates may be obtained by passing sulphuretted hydrogen through solutions of arsenates:—

$$\underset{\text{Hydric disodic arsenate.}}{AsOHoNao_2} + \underset{\text{Sulphuretted hydrogen.}}{4SH_2} = \underset{\text{Sulphhydric disulphosodic sulpharsenate.}}{AsSHsNas_2} + \underset{\text{Water.}}{4OH_2}.$$

ANTIMONY, Sb_4?

Atomic weight =122. *Probable molecular weight* =488. *Sp. gr. crystalline* 6·7, *amorphous* 5·78. *Fuses at* 430°. *Atomicity* $'''$ *and* $^{\text{v}}$. *Evidence of atomicity*:—

Antimonious chloride	$Sb'''Cl_3$.
Antimonic tetretho-chloride (*Tetrethyl-stibonic chloride*).	$Sb^{\text{v}}Et_4Cl$.
Antimonic chloride	$Sb^{\text{v}}Cl_5$.

Occurrence.—To a small extent in the native state. Alloyed with metals in a few minerals. Sometimes in the form of oxide, but principally in the form of grey antimony ore or ***stibnite***, which consists of antimonious sulphide.

a. *Crystalline Antimony.*

Preparation.—1. By fusing the native sulphide and introducing metallic iron, which combines with the sulphur:—

$$\underset{\text{Antimonious sulphide.}}{Sb_2S''_3} + 3Fe = \underset{\text{Ferrous sulphide.}}{3FeS''} + Sb_2.$$

2. The native sulphide is roasted in contact with the air, when it is partially converted into antimonious oxide:—

$$\underset{\text{Antimonious sulphide.}}{2Sb_2S''_3} + 9O_2 = \underset{\text{Antimonious oxide.}}{2Sb_2O_3} + \underset{\text{Sulphurous anhydride.}}{6SO_2}.$$

The roasted mineral is then fused with charcoal and sodic carbonate. The reaction takes place in two stages: first, the remaining sulphide is converted into oxide by the sodic carbonate, and subsequently the oxide is reduced by the carbon:—

$$1.\ \underset{\text{Antimonious sulphide.}}{Sb_2S''_3} + \underset{\text{Sodic carbonate.}}{3CONao_2} = \underset{\text{Carbonic anhydride.}}{3CO_2} + \underset{\text{Sodic sulphide.}}{3SNa_2} + \underset{\text{Antimonious oxide.}}{Sb_2O_3};$$

$$2.\ \underset{\text{Antimonious oxide.}}{Sb_2O_3} + 3C = \underset{\text{Carbonic oxide.}}{3CO} + Sb_2.$$

3. Antimony may be obtained in the pure condition by reducing, with charcoal, the oxide formed by the action of nitric acid upon crude antimony.

β. *Amorphous Antimony.*

Preparation.—By the electrolysis of a solution of tartar emetic in antimonious chloride.

COMPOUND OF ANTIMONY WITH HYDROGEN.

ANTIMONIURETTED HYDROGEN, *Antimonious Hydride.*

$$SbH_3.$$

Molecular weight =125.

This compound is unknown in the pure condition.

Preparation.—1. By the action of hydrochloric acid upon an alloy of zinc and antimony:—

$$\underset{\text{Antimonious zincide.}}{Sb_2Zn''_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Antimonious hydride.}}{2SbH_3} + \underset{\text{Zincic chloride.}}{3ZnCl_2}.$$

2. By the action of nascent hydrogen evolved from zinc and sulphuric acid upon soluble antimony compounds. In both these reactions the antimonious hydride is always mixed with much hydrogen:—

$$\underset{\text{Antimonious chloride.}}{SbCl_3} + 3H_2 = \underset{\text{Antimonious hydride.}}{SbH_3} + \underset{\text{Hydrochloric acid.}}{3HCl}.$$

Reactions.—1. When burnt in air or in oxygen, it yields water and antimonious oxide:—

$$\underset{\text{Antimonious hydride.}}{2SbH_3} + 3O_2 = \underset{\text{Antimonious oxide.}}{Sb_2O_3} + \underset{\text{Water.}}{3OH_2}.$$

2. When burnt with a limited supply of air the hydrogen only is oxidized, the antimony being deposited:—

$$\underset{\text{Antimonious hydride.}}{4SbH_3} + 3O_2 = Sb_4 + \underset{\text{Water.}}{6OH_2}.$$

3. Decomposed into its elements, like arsenious hydride, when passed through a red-hot tube.

4. When transmitted through a solution of argentic nitrate, it produces a precipitate of antimonious argentide, thus differing from arsenious hydride (see p. 125):—

$$\underset{\text{Argentic nitrate.}}{3NO_2Ago} + \underset{\text{Antimonious hydride.}}{SbH_3} = \underset{\text{Nitric acid.}}{3NO_2Ho} + \underset{\text{Antimonious argentide.}}{SbAg_3}.$$

From the composition of this compound, and from that of some of its analogues, the composition of antimonious hydride is inferred.

Antimonious hydride	$\mathbf{Sb}H_3$.
Antimonious bromide	$\mathbf{Sb}Br_3$.
Antimonious argentide	$\mathbf{Sb}Ag_3$.
Antimonious zincide	$\mathbf{Sb}_2Zn''_3$.
Antimonious ethide. (*Triethylstibine.*)	$\mathbf{Sb}Et_3$.
Antimonious amylide. (*Triamylstibine.*)	$\mathbf{Sb}Ay_3$.

COMPOUNDS OF ANTIMONY WITH CHLORINE.

Antimonious chloride	$\mathbf{Sb}Cl_3$.
Antimonic chloride	$\mathbf{Sb}Cl_5$.

ANTIMONIOUS CHLORIDE.

$\mathbf{Sb}Cl_3$.

Molecular weight = 228·5. *Molecular volume* $\boxed{}\!\boxed{}$. 1 *litre of antimonious chloride vapour weighs* 114·25 *criths. Fuses at* 72°. *Boils at* 223°.

Preparation.—1. By passing chlorine over excess of metallic antimony or antimonious sulphide, and purifying by distillation:—

$$Sb_2 + 3Cl_2 = \underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3}.$$

$$\underset{\text{Antimonious sulphide.}}{2\mathbf{Sb}_2S''_3} + 9Cl_2 = \underset{\text{Antimonious chloride.}}{4\mathbf{Sb}Cl_3} + \underset{\text{Disulphur dichloride.}}{3'\mathbf{S}'_2Cl_2}.$$

2. By dissolving antimonious sulphide in hydrochloric acid, or antimony in hydrochloric acid containing a little nitric acid, evaporating and distilling the product:—

$$\underset{\text{Antimonious sulphide.}}{\mathbf{Sb}_2S''_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Sulphuretted hydrogen.}}{3\mathbf{S}H_2} + \underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3}.$$

$$Sb_2 + \underset{\text{Hydrochloric acid.}}{6HCl} + \underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho} = \underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3} + \underset{\text{Water.}}{6\mathbf{O}H_2} + \underset{\text{Nitric peroxide.}}{3\mathbf{N}^{iv}_2O_4}.$$

3. By distilling antimony or antimonious sulphide with mercuric chloride:—

$$Sb_4 + \underset{\text{Mercuric chloride.}}{4\mathbf{Hg}Cl_2} = \underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3} + \underset{\text{Dimercuric diantimonide.}}{'\mathbf{Sb}''_2Hg''_3} + \underset{\text{Mercurous chloride.}}{'\mathbf{Hg}'_2Cl_2}.$$

$$\underset{\text{Antimonious sulphide.}}{\mathbf{Sb}_2S''_3} + \underset{\text{Mercuric chloride.}}{3\mathbf{Hg}Cl_2} = \underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3} + \underset{\text{Mercuric sulphide.}}{3\mathbf{Hg}S''}.$$

4. By distilling antimonious sulphate with sodic chloride:—

$$\underset{\text{Antimonious sulphate.}}{\mathbf{S}_3O_6Sbo'''_2} + \underset{\text{Sodic chloride.}}{6NaCl} = \underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3} + \underset{\text{Sodic sulphate.}}{3\mathbf{S}O_2Nao_2}.$$

Reaction.—With water it produces antimonious oxychloride:—

$$\underset{\text{Antimonious chloride.}}{\mathbf{Sb}Cl_3} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Antimonious oxychloride.}}{\mathbf{Sb}OCl}.$$

Long-continued action of water transforms this compound into antimonious oxide:—

$$\underset{\text{Antimonious oxychloride.}}{2\mathbf{Sb}OCl} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Antimonious oxide.}}{\mathbf{Sb}_2O_3}.$$

ANTIMONIC CHLORIDE.

$\mathbf{Sb}Cl_5$.

Molecular weight **=299·5.** ***Fuses at*** 0°.

Preparation.—1. By acting upon antimony with excess of chlorine:—

$$Sb_2 + 5Cl_2 = \underset{\text{Antimonic chloride.}}{2\mathbf{Sb}Cl_5}.$$

2. By passing chlorine over antimonious chloride, the latter liquefies, producing antimonic chloride :—

$$\underset{\text{Antimonious chloride.}}{\mathbf{Sb}Cl_3} + Cl_2 = \underset{\text{Antimonic chloride.}}{\mathbf{Sb}Cl_5}.$$

Reactions.—1. With a small quantity of water it forms antimonic oxytrichloride, analogous to phosphoric oxytrichloride :—

$$\underset{\text{Antimonic chloride.}}{\mathbf{Sb}Cl_5} + \underset{\text{Water.}}{OH_2} = \underset{\text{Antimonic oxytrichloride.}}{\mathbf{Sb}OCl_3} + \underset{\text{Hydrochloric acid.}}{2HCl}.$$

2. An excess of water transforms antimonic chloride into orthantimonic acid or pyrantimonic acid, corresponding to pyrophosphoric acid :—

$$\underset{\text{Antimonic chloride.}}{\mathbf{Sb}Cl_5} + \underset{\text{Water.}}{4OH_2} = \underset{\text{Orthantimonic acid.}}{\mathbf{Sb}OHo_3} + \underset{\text{Hydrochloric acid.}}{5HCl}:$$

$$\text{or } \underset{\text{Antimonic chloride.}}{2\mathbf{Sb}Cl_5} + \underset{\text{Water.}}{7OH_2} = \underset{\text{Pyrantimonic acid.}}{\mathbf{Sb}_2O_3Ho_4} + \underset{\text{Hydrochloric acid.}}{10HCl}.$$

3. By the action of sulphuretted hydrogen antimonic sulphotrichloride is formed :—

$$\underset{\text{Antimonic chloride.}}{\mathbf{Sb}Cl_5} + \underset{\text{Sulphuretted hydrogen.}}{\mathbf{S}H_2} = \underset{\text{Antimonic sulphotrichloride.}}{\mathbf{Sb}S''Cl_3} + \underset{\text{Hydrochloric acid.}}{2HCl}.$$

Antimonious bromide, $\mathbf{Sb}Br_3$, resembles antimonious chloride; it fuses at 90°, boils at 270°, and by the action of water is converted into the *oxybromide*, $\mathbf{Sb}OBr$.

Antimonious iodide, $\mathbf{Sb}I_3$, when acted upon by water forms the *oxyiodide*, $\mathbf{Sb}OI$.

The corresponding *fluoride*, $\mathbf{Sb}F_3$, is said to exist and to be soluble in water without decomposition.

OXIDES AND ACIDS OF ANTIMONY.

Antimonious oxide or anhydride	$\mathbf{Sb}_2O_3$.
Diantimonic tetroxide	$'\mathbf{Sb}^{iv}{}_2O_4$.
Antimonic anhydride	$\mathbf{Sb}_2O_5$.
Metantimonious acid	$\mathbf{Sb}OHo$.
Orthantimonic acid	$\mathbf{Sb}OHo_3$?
Metantimonic acid	$\mathbf{Sb}O_2Ho$.
Pyrantimonic acid	$\mathbf{Sb}_2O_3Ho_4$.

ANTIMONIOUS OXIDE, OR ANHYDRIDE.

$$\mathbf{Sb}_2O_3.$$

Molecular weight = 292.

Occurrence.—In nature in the rare minerals *valentinite* and *senarmontite.*

Preparation.—1. By burning antimony in air:—

$$2Sb_2 + 3O_2 = \underset{\text{Antimonious oxide.}}{2\mathbf{Sb}_2O_3}.$$

2. By pouring a solution of antimonious chloride in dilute hydrochloric acid into a boiling solution of sodic carbonate:—

$$\underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3} + \underset{\text{Sodic carbonate.}}{3\mathbf{C}ONao_2} = \underset{\text{Antimonious oxide.}}{\mathbf{Sb}_2O_3} + \underset{\text{Carbonic anhydride.}}{3\mathbf{C}O_2} + \underset{\text{Sodic chloride.}}{6NaCl}.$$

3. By heating metantimonious acid to the temperature of boiling water:—

$$\underset{\text{Metantimonious acid.}}{2\mathbf{Sb}OHo} = \underset{\text{Water.}}{\mathbf{O}H_2} + \underset{\text{Antimonious oxide.}}{\mathbf{Sb}_2O_3}.$$

Reactions.—1. When heated to redness in the air, it burns like tinder, forming diantimonic tetroxide:—

$$\underset{\text{Antimonious oxide.}}{\mathbf{Sb}_2O_3} + O = \underset{\text{Diantimonic tetroxide.}}{'\mathbf{Sb}^{iv}{}_2O_4}.$$

2. Readily reduced to the metallic state by ignition with charcoal, hydrogen, &c.

$$\underset{\text{Antimonious oxide.}}{Sb_2O_3} + C_3 = Sb_2 + \underset{\text{Carbonic oxide.}}{3CO}.$$

$$\underset{\text{Antimonious oxide.}}{Sb_2O_3} + 3H_2 = Sb_2 + \underset{\text{Water.}}{3OH_2}.$$

3. Readily dissolved by a hot solution of hydric potassic tartrate (cream of tartar), forming potassic antimonylic tartrate (tartar emetic) :—

$$\underset{\text{Hydric potassic tartrate (Cream of tartar).}}{2\left\{\begin{array}{l}COHo\\CHHo\\CHHo\\COKo\end{array}\right.} + \underset{\text{Antimonious oxide.}}{Sb_2O_3} = \underset{\text{Potassic antimonylic tartrate (Tartar emetic).}}{2\left\{\begin{array}{l}CO(Sb'''O_2)\\CHHo\\CHHo\\COKo\end{array}\right.} + \underset{\text{Water.}}{OH_2}.$$

4. Dissolved by hydrochloric acid, forming antimonious chloride :—

$$\underset{\text{Antimonious oxide.}}{Sb_2O_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Antimonious chloride.}}{2SbCl_3} + \underset{\text{Water.}}{3OH_2}.$$

METANTIMONIOUS ACID.

$SbOHo.$

Molecular weight = 155.

Preparation.—By pouring a solution of antimonious chloride into a cold solution of sodic carbonate :—

$$\underset{\text{Antimonious chloride.}}{2SbCl_3} + \underset{\text{Sodic carbonate.}}{3CONao_2} + \underset{\text{Water.}}{OH_2} = \underset{\text{Metantimonious acid.}}{2SbOHo}$$

$$+ \underset{\text{Carbonic anhydride.}}{3CO_2} + \underset{\text{Sodic chloride.}}{6NaCl}.$$

Reactions.—1. Decomposed by heat (page 186).

2. Readily dissolved by alkaline hydrates, producing ill-defined antimonites.

DIANTIMONIC TETROXIDE.

$$\left\{\begin{matrix}\mathbf{Sb}O_2\\ \mathbf{Sb}O_2\end{matrix}\right. = {}'\mathbf{Sb}^{iv}{}_2O_4.$$

Molecular weight =308.

Occurrence.—Found native as *cervantite.*

Preparation.—1. By igniting antimonic oxide, or the white solid produced by the action of nitric acid upon metallic antimony:—

$$\underset{\text{Antimonic oxide.}}{2\mathbf{Sb}_2O_5} = \underset{\text{Diantimonic tetroxide.}}{2'\mathbf{Sb}^{iv}{}_2O_4} + O_2.$$

2. By heating antimonious oxide in contact with the air:—

$$\underset{\text{Antimonious oxide.}}{2\mathbf{Sb}_2O_3} + O_2 = \underset{\text{Diantimonic tetroxide.}}{2'\mathbf{Sb}^{iv}{}_2O_4}.$$

ANTIMONIC ANHYDRIDE.

$$\mathbf{Sb}_2O_5.$$

Molecular weight =324. *Sp. gr.* 6·6.

Preparation.—By gently heating the corresponding acids:—

$$\underset{\text{Orthantimonic acid.}}{2\mathbf{Sb}O\mathrm{Ho}_3} = \underset{\text{Antimonic anhydride.}}{\mathbf{Sb}_2O_5} + \underset{\text{Water.}}{3OH_2}.$$

$$\underset{\text{Metantimonic acid.}}{2\mathbf{Sb}O_2\mathrm{Ho}} = \underset{\text{Antimonic anhydride.}}{\mathbf{Sb}_2O_5} + \underset{\text{Water.}}{OH_2}.$$

$$\underset{\substack{\text{Pyrantimonic}\\ \text{acid.}}}{Sb_2O_3Ho_4} = \underset{\substack{\text{Antimonic}\\ \text{anhydride.}}}{Sb_2O_5} + \underset{\text{Water.}}{2OH_2}.$$

Reactions.—1. When heated, it is decomposed into diantimonic tetroxide and oxygen (see page 138).

2. Fused with potassic carbonate, it produces potassic metantimoniate :—

$$\underset{\substack{\text{Antimonic}\\ \text{anhydride.}}}{Sb_2O_5} + \underset{\substack{\text{Potassic}\\ \text{carbonate.}}}{COKo_2} = \underset{\substack{\text{Potassic}\\ \text{metantimonate.}}}{2SbO_2Ko} + \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{CO_2}.$$

ORTHANTIMONIC ACID ?

$SbOHo_3$?

Preparation.—Said to be formed by the action of water upon antimonic chloride (see p. 135).

METANTIMONIC ACID.

$SbO_2Ho.$

Preparation.—1. By the action of nitric acid containing a little hydrochloric acid on metallic antimony :—

$$Sb_2 + \underset{\text{Nitric acid.}}{4NO_2Ho} = \underset{\substack{\text{Metantimonic}\\ \text{acid.}}}{2SbO_2Ho} + \underset{\substack{\text{Nitrous}\\ \text{anhydride.}}}{N_2O_3}$$

$$+ \underset{\text{Nitric oxide.}}{N''_2O_2} + \underset{\text{Water.}}{OH_2}.$$

2. By the spontaneous dehydration of orthantimonic acid, or of pyrantimonic acid :—

$$\underset{\substack{\text{Orthantimonic}\\ \text{acid.}}}{SbOHo_3} = \underset{\text{Water.}}{OH_2} + \underset{\substack{\text{Metantimonic}\\ \text{acid.}}}{SbO_2Ho};$$

$$\underset{\substack{\text{Pyrantimonic}\\ \text{acid.}}}{Sb_2O_3Ho_4} = \underset{\text{Water.}}{OH_2} + \underset{\substack{\text{Metantimonic}\\ \text{acid.}}}{2SbO_2Ho}.$$

Reaction.—By the action of alkaline hydrates it produces either metantimonates or orthantimonates:—

$$\underset{\text{Metantimonic acid.}}{\mathbf{Sb}O_2Ho} + \underset{\text{Potassic hydrate.}}{\mathbf{O}KH} = \underset{\text{Potassic metantimonate.}}{\mathbf{Sb}O_2Ko} + \underset{\text{Water.}}{\mathbf{O}H_2};$$

$$\underset{\text{Metantimonic acid.}}{\mathbf{Sb}O_2Ho} + \underset{\text{Potassic hydrate.}}{\mathbf{O}KH} = \underset{\text{Dihydric potassic orthantimonate.}}{\mathbf{Sb}OHo_2Ko.}$$

PYRANTIMONIC ACID, *Parantimonic Acid.*

(*Metantimonic acid of Frémy.*)

$\mathbf{Sb}_2O_3Ho_4$.

Preparation.—By acidifying solutions of pyrantimonates:—

$$\underset{\text{Dihydric dipotassic pyrantimonate.}}{\mathbf{Sb}_2O_3Ho_2Ko_2} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Pyrantimonic acid.}}{\mathbf{Sb}_2O_3Ho_4} + \underset{\text{Potassic chloride.}}{2KCl.}$$

Dihydric dipotassic pyrantimonate is prepared by fusing antimonic anhydride with excess of potassic hydrate, and extracting the mass with water, when an alkaline solution containing dihydric dipotassic pyrantimonate $\mathbf{Sb}_2O_3Ho_2Ko_2$ is formed. This solution produces precipitates in solutions of sodium salts, the sodic pyrantimonate thus formed containing $\mathbf{Sb}_2O_3Ho_2Nao_2, 6\mathbf{O}H_2$.

COMPOUND OF ANTIMONY WITH OXYGEN AND SULPHUR.

ANTIMONIOUS OXYDISULPHIDE.

$$\left\{\begin{array}{l}\mathbf{Sb}S'' \\ O \\ \mathbf{Sb}S''\end{array}\right. = \mathbf{Sb}_2S''_2O.$$

Molecular weight = 324.

Occurs as a rare mineral known as *red antimony*.

COMPOUNDS OF ANTIMONY AND SULPHUR.

Antimonious sulphide .... $\mathbf{Sb}_2S''_3$.
Antimonic sulphide .... $\mathbf{Sb}_2S''_5$.

ANTIMONIOUS SULPHIDE, *Sulphantimonious Anhydride.*

$$\mathbf{Sb}_2S''_3.$$

Molecular weight = 340.

Occurrence.—In nature as *stibnite* or *grey antimony ore.*

Preparation.—1. By heating together antimony and sulphur, or antimonious oxide and sulphur, in the proper proportions :—

$$2Sb_2 + 3S_2 = \underset{\text{Antimonious sulphide.}}{2\mathbf{Sb}_2S''_3}.$$

$$\underset{\text{Antimonious oxide.}}{2\mathbf{Sb}_2O_3} + S_9 = \underset{\text{Antimonious sulphide.}}{2\mathbf{Sb}_2S''_3} + \underset{\text{Sulphurous anhydride.}}{3\mathbf{S}O_2}.$$

2. By passing sulphuretted hydrogen through a solution of antimonious chloride :—

$$\underset{\text{Antimonious chloride.}}{2\mathbf{Sb}Cl_3} + \underset{\text{Sulphuretted hydrogen.}}{3\mathbf{S}H_2} = \underset{\text{Antimonious sulphide.}}{\mathbf{Sb}_2S''_3} + \underset{\text{Hydrochloric acid.}}{6HCl}.$$

Reactions.—1. Decomposed by hot hydrochloric acid (see p. 133).

2. Soluble with decomposition in solutions of alkaline hydrates :—

$$\underset{\text{Antimonious sulphide.}}{\mathbf{Sb}_2S''_3} + \underset{\text{Potassic hydrate.}}{6KHo} = \underset{\text{Trisulphopotassic sulphantimonite.}}{\mathbf{Sb}Ks_3} + \underset{\text{Tripotassic antimonite.}}{\mathbf{Sb}Ko_3} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

Addition of an acid reproduces and precipitates the antimonious sulphide :—

$$\underset{\text{Trisulphopotassic sulphantimonite.}}{\mathbf{Sb}Ks_3} + \underset{\text{Tripotassic antimonite.}}{\mathbf{Sb}Ko_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Antimonious sulphide.}}{\mathbf{Sb}_2S''_3} + \underset{\text{Potassic chloride.}}{6KCl} + \underset{\text{Water.}}{3\mathbf{O}H_2}.$$

3. Soluble in alkaline sulphhydrates:—

$$\begin{array}{ccccccc} \mathbf{Sb}_2S''_3 & + & 6KHs & = & 2\mathbf{Sb}Ks_3 & + & 3SH_2. \\ \text{Antimonious sulphide.} & & \text{Potassic sulphhydrate.} & & \text{Trisulphopotassic sulphantimonite.} & & \text{Sulphuretted hydrogen.} \end{array}$$

SULPHANTIMONITES.

Many sulphantimonites occur in nature:—

Orthosulphantimonites.

General formulæ:—$\mathbf{Sb}Ms_3$ and $\mathbf{Sb}_2Ms''_3$.

Dark-red silver. *Trisulphargentic sulphantimonite*	$\mathbf{Sb}Ags_3$.
Boulangerite. *Trisulphoplumbic sulphantimonite*	$\mathbf{Sb}_2Pbs'_3$.
Bournonite. *Disulphoplumbic sulphocuprous sulphantimonite*	$\mathbf{Sb}_2Pbs''_2(Cu_2S''_2)''$.

Metasulphantimonites.

General formulæ:—$\mathbf{Sb}S''Ms$ and $\mathbf{Sb}_2S''_2Ms''$.

Miargyrite. *Sulphargentic metasulphantimonite*	$\mathbf{Sb}S''Ags$.
Zinkenite. *Sulphoplumbic metasulphantimonite*	$\mathbf{Sb}_2S''_2Pbs''$.
Antimony copper glance. *Sulphocuprous metasulphantimonite*	$\mathbf{Sb}_2S''_2(Cu_2S''_2)''$.
Berthierite. *Sulphoferrous metasulphantimonite*	$\mathbf{Sb}_2S''_2Fes''$.

Pyrosulphantimonites.

General formulæ:—$\mathbf{Sb}_2S''Ms_4$ and $\mathbf{Sb}_2S''Ms''_2$.

Feather ore. *Sulphoplumbic pyrosulphantimonite*	$\mathbf{Sb}_2S''Pbs''_2$.
Fahl ore. *Sulphocuprosoferrous pyrosulphantimonite*	$\mathbf{Sb}_2S''(Cu_2FeS''_2)''_2$.

ANTIMONIC SULPHIDE, *Sulphantimonic Anhydride.*

$Sb_2S''_5$.

Molecular weight = 404.

Preparation.—1. By passing sulphuretted hydrogen through a solution of antimonic chloride:—

$$\underset{\text{Antimonic chloride.}}{2SbCl_5} + \underset{\text{Sulphuretted hydrogen.}}{5SH_2} = \underset{\text{Antimonic sulphide.}}{Sb_2S''_5} + \underset{\text{Hydrochloric acid.}}{10HCl}.$$

2. By the addition of an acid to a solution of a sulphantimonate:—

$$\underset{\text{Trisulphosodic sulphantimonate.}}{2SbS''Nas_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Antimonic sulphide.}}{Sb_2S''_5} + \underset{\text{Sodic chloride.}}{6NaCl} + \underset{\text{Sulphuretted hydrogen.}}{3SH_2}.$$

Reactions.—1. Decomposed by boiling hydrochloric acid, into antimonious chloride, sulphuretted hydrogen, and sulphur:—

$$\underset{\text{Antimonic sulphide.}}{Sb_2S''_5} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Antimonious chloride.}}{2SbCl_3} + \underset{\text{Sulphuretted}}{3SH_2} + S_2.$$

2. Soluble in solutions of alkaline sulphides:—

$$\underset{\text{Antimonic sulphide.}}{Sb_2S''_5} + \underset{\text{Potassic sulphide.}}{3SK_2} = \underset{\text{Trisulphopotassic sulphantimonate.}}{2SbS''Ks_3}.$$

3. Soluble in solutions of alkaline hydrates:—

$$\underset{\text{Antimonic sulphide.}}{4Sb_2S''_5} + \underset{\text{Potassic hydrate.}}{24OKH} = \underset{\text{Tripotassic antimonate.}}{3SbOKo_3} + \underset{\text{Trisulphopotassic sulphantimonate.}}{5SbS''Ks_3} + \underset{\text{Water.}}{12OH_2}.$$

BISMUTH, Bi_4?

Atomic weight =208. *Sp. gr.* 9·83. *Fuses at* 265°. *Atomicity* $'''$ *and* ^{v}. *Evidence of atomicity*:—

Bismuthous chloride	$\mathbf{Bi}'''Cl_3$.
Bismuthous oxide	$\mathbf{Bi}'''_2O_3$.
Bismuthous ethide	$\mathbf{Bi}'''Et_3$.
Bismuthous dichlorethide	$\mathbf{Bi}'''EtCl_2$.
Bismuthic anhydride	$\mathbf{Bi}^{v}_2O_5$.

Occurrence.—Principally in the metallic state in nature.

Preparation.—1. On a large scale by fusion and separation from earthy impurities.

2. It may be obtained in the pure state by dissolving commercial bismuth in nitric acid, precipitating the basic nitrate by addition of water, and reducing the precipitate by ignition with charcoal.

No compound of bismuth with hydrogen is known.

COMPOUND OF BISMUTH WITH CHLORINE.

BISMUTHOUS CHLORIDE.

$\mathbf{Bi}Cl_3$.

Molecular weight =314·5. *Molecular volume* $\square\square$. 1 *litre of bismuthous chloride vapour weighs* 157·25 *criths.*

Preparation.—1. By passing dry chlorine over metallic bismuth:—

$$Bi_2 \quad + \quad 3Cl_2 \quad = \quad \underset{\text{Bismuthous chloride.}}{2\mathbf{Bi}Cl_3.}$$

2. By evaporating a solution of bismuth in hydrochloric acid containing a little nitric acid and distilling.

3. By distilling metallic bismuth with mercuric chloride :—

$$\mathbf{Bi}_2 + \underset{\text{Mercuric chloride.}}{6\mathbf{Hg}Cl_2} = \underset{\text{Bismuthous chloride.}}{2\mathbf{Bi}Cl_3} + \underset{\text{Mercurous chloride.}}{3\mathbf{Hg}'_2Cl_2}.$$

Reaction.—By the addition of water it is decomposed, forming bismuthous oxychloride :—

$$\underset{\text{Bismuthous chloride.}}{\mathbf{Bi}Cl_3} + \underset{\text{Water.}}{OH_2} = \underset{\text{Bismuthous oxychloride.}}{\mathbf{Bi}OCl} + \underset{\text{Hydrochloric acid.}}{2HCl}.$$

The following compounds are also known :—

Bismuthous bromide	$\mathbf{Bi}Br_3$.
Bismuthous iodide	$\mathbf{Bi}I_3$.
Bismuthous fluoride	$\mathbf{Bi}F_3$.
Bismuthous oxybromide	$\mathbf{Bi}OBr$
Bismuthous oxyiodide	$\mathbf{Bi}OI$.
Dibismuthous tetrachloride	$\left\{\begin{matrix}\mathbf{Bi}Cl_2.\\ \mathbf{Bi}Cl_2.\end{matrix}\right.$

COMPOUNDS OF BISMUTH WITH OXYGEN AND HYDROXYL.

Dibismuthous dioxide	$\left\{\begin{matrix}\mathbf{Bi}O.\\ \mathbf{Bi}O.\end{matrix}\right.$
Bismuthous oxide	$\mathbf{Bi}_2O_3$.
Dibismuthic tetroxide	$\mathbf{Bi}^{iv}_2O_4$.
Bismuthic anhydride	$\mathbf{Bi}_2O_5$.
Bismuthous oxyhydrate, or metabismuthous acid	$\mathbf{Bi}OHo$.
Metabismuthic acid	$\mathbf{Bi}O_2Ho$

BISMUTHOUS OXIDE.

$\mathbf{Bi}_2O_3$.

Molecular weight =464. *Sp. gr.* 8·2.

Occurrence.—As the rare mineral ***bismuth ochre.***

Preparation.—1. By burning bismuth in air or oxygen.

2. By heating the nitrate, carbonate, or hydrate:—

$$\underset{\text{Bismuthous nitrate.}}{2\mathbf{N}_3O_6Bio'''} = \underset{\text{Bismuthous oxide.}}{\mathbf{Bi}_2O_3} + \underset{\text{Nitrous anhydride.}}{3\mathbf{N}_2O_3} + 3O_2.$$

$$\underset{\text{Bismuthous nitrate dihydrate.}}{2\mathbf{N}O_2(Bi'''Ho_2O)} = \underset{\text{Bismuthous oxide.}}{\mathbf{Bi}_2O_3} + \underset{\text{Water.}}{2\mathbf{O}H_2} + \underset{\text{Nitrous anhydride.}}{\mathbf{N}_2O_3} + O_2.$$

$$\underset{\text{Bismuthylic carbonate.}}{\mathbf{C}O(BiO_2)_2} = \underset{\text{Bismuthous oxide.}}{\mathbf{Bi}_2O_3} + \underset{\text{Carbonic anhydride.}}{\mathbf{C}O_2}.$$

$$\underset{\text{Bismuthous oxyhydrate.}}{2\mathbf{Bi}OHo} = \underset{\text{Bismuthous oxide.}}{\mathbf{Bi}_2O_3} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

3. By boiling bismuthous hydrate in solution of potassic hydrate, whereby it is converted into bismuthous oxide with loss of water.

Reactions.—Dissolved by hydrochloric, nitric, and sulphuric acids, forming the bismuthous chloride, nitrate, and sulphate:—

$$\underset{\text{Bismuthous chloride.}}{\mathbf{Bi}Cl_3.} \qquad \underset{\text{Bismuthous nitrate.}}{\mathbf{N}_3O_6Bio'''.} \qquad \underset{\text{Bismuthous sulphate.}}{\mathbf{S}_3O_6Bio'''_2.}$$

BISMUTHOUS OXYHYDRATE, *Metabismuthous Acid.*

$\mathbf{Bi}OHo$.

Preparation.—By pouring a solution of bismuthous nitrat in dilute nitric acid into dilute ammonia or potassic hydrate, and drying the precipitate, which, at first, probably contains orthobismuthous acid:—

$$\underset{\substack{\text{Bismuthous}\\\text{nitrate.}}}{\mathbf{N}_3O_6Bio'''} + \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{3\mathbf{O}KH} = \underset{\substack{\text{Orthobismuthous}\\\text{acid.}}}{\mathbf{Bi}Ho_3} + \underset{\substack{\text{Potassic}\\\text{nitrate.}}}{3\mathbf{N}O_2Ko.}$$

$$\underset{\substack{\text{Orthobismuthous}\\\text{acid.}}}{\mathbf{Bi}Ho_3} = \underset{\substack{\text{Metabismuthous}\\\text{acid.}}}{\mathbf{Bi}OHo} + \underset{\text{Water.}}{\mathbf{O}H_2.}$$

Reaction.—By heat or by boiling with caustic alkali, water is expelled, and bismuthous oxide formed (see p. 146).

An unstable metabismuthite is produced by fusing bismuthous oxide with sodic carbonate:—

$$\underset{\substack{\text{Bismuthous}\\\text{oxide.}}}{\mathbf{Bi}_2O_3} + \underset{\substack{\text{Sodic}\\\text{carbonate.}}}{\mathbf{C}ONao_2} = \underset{\substack{\text{Sodic}\\\text{metabismuthite.}}}{2\mathbf{Bi}ONao} + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{\mathbf{C}O_2.}$$

BISMUTHIC ANHYDRIDE.

$$\mathbf{Bi}_2O_5.$$

Preparation.—By heating bismuthic acid to 130°.

Reactions.—1. When heated to the boiling-point of mercury, it loses oxygen, being converted either into bismuthous oxide or dibismuthic tetroxide:—

$$\underset{\substack{\text{Bismuthic}\\\text{anhydride.}}}{\mathbf{Bi}_2O_5} = \underset{\substack{\text{Bismuthous}\\\text{oxide.}}}{\mathbf{Bi}_2O_3} + O_2:$$

$$\underset{\substack{\text{Bismuthic}\\\text{anhydride.}}}{2\mathbf{Bi}_2O_5} = \underset{\substack{\text{Dibismuthic}\\\text{tetroxide.}}}{2{}'\mathbf{Bi}^{iv}{}_2O_4} + O_2.$$

2. When heated in a current of hydrogen, it is readily reduced to bismuthous oxide.

3. Heated with hydrochloric acid, it evolves chlorine, producing bismuthous chloride and water :—

$$\underset{\text{Bismuthic anhydride.}}{\mathbf{Bi}_2O_5} + \underset{\text{Hydrochloric acid.}}{10HCl} = \underset{\text{Bismuthous chloride,}}{2\mathbf{Bi}Cl_3} + \underset{\text{Water.}}{5\mathbf{O}H_2} + 2Cl_2.$$

4. Sulphurous acid converts it into bismuthous sulphate :—

$$\underset{\text{Bismuthic anhydride.}}{3\mathbf{Bi}_2O_5} + \underset{\text{Sulphurous acid.}}{6\mathbf{S}OHo_2} = \underset{\text{Bismuthous sulphate.}}{2\mathbf{S}_3O_6Bio'''_2} + \underset{\text{Bismuthous oxide.}}{\mathbf{Bi}_2O_3} + \underset{\text{Water.}}{6\mathbf{O}H_2}.$$

5. When heated with sulphuric or nitric acid, it evolves oxygen, producing bismuthous sulphate or nitrate :—

$$\underset{\text{Bismuthic anhydride.}}{\mathbf{Bi}_2O_5} + \underset{\text{Sulphuric acid.}}{3\mathbf{S}O_2Ho_2} = \underset{\text{Bismuthous sulphate.}}{\mathbf{S}_3O_6Bio'''_2} + \underset{\text{Water.}}{3\mathbf{O}H_2} + O_2;$$

$$\underset{\text{Bismuthic anhydride.}}{\mathbf{Bi}_2O_5} + \underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho} = \underset{\text{Bismuthous nitrate.}}{2\mathbf{N}_3O_6Bio'''} + \underset{\text{Water.}}{3\mathbf{O}H_2} + O_2.$$

METABISMUTHIC ACID.

$\mathbf{Bi}O_2Ho.$

Preparation.—Obtained as a red deposit by passing chlorine through a solution of potassic hydrate containing bismuthous oxide in suspension :—

$$\underset{\text{Potassic hydrate.}}{4\mathbf{O}KH} + 2Cl_2 + \underset{\text{Bismuthous oxide.}}{\mathbf{Bi}_2O_3} = \underset{\text{Metabismuthic acid.}}{2\mathbf{Bi}O_2Ho} + \underset{\text{Potassic chloride.}}{4KCl} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

Reaction.—Dissolves in hot solution of potassic hydrate. By the addition of an acid to the liquid, a salt, said to have the composition

$$\mathbf{Bi}_2O_4HoKo,$$

is precipitated.

COMPOUND OF BISMUTH WITH SULPHUR.

Dibismuthous disulphide ...... $'\mathbf{Bi}''_2S''_2$.
Bismuthous sulphide .... $\mathbf{Bi}_2S''_3$.

DIBISMUTHOUS DISULPHIDE.

$$\left\{\begin{matrix}\mathbf{Bi}S''\\ \mathbf{Bi}S''\end{matrix}\right., \text{ or } '\mathbf{Bi}''_2S''_2.$$

Molecular weight =480. *Sp. gr.* 7·3.

Preparation.—By fusing bismuth and sulphur together in the proper proportions.

BISMUTHOUS SULPHIDE.

$$\mathbf{Bi}_2S''_3.$$

Molecular weight =512. *Sp. gr.* 6·4.

Occurrence.—As the rare mineral *bismuth glance.*

Preparation.—1. By fusing sulphur and bismuth in the proper proportions.

2. By precipitating bismuth solutions by sulphuretted hydrogen:—

$$\underset{\substack{\text{Bismuthous}\\ \text{chloride.}}}{2\mathbf{Bi}Cl_3} + \underset{\substack{\text{Sulphuretted}\\ \text{hydrogen.}}}{3\mathbf{S}H_2} = \underset{\substack{\text{Bismuthous}\\ \text{sulphide.}}}{\mathbf{Bi}_2S''_3} + \underset{\substack{\text{Hydrochloric}\\ \text{acid.}}}{6HCl.}$$

Reaction.—This compound is not dissolved by alkaline hydrates or sulphhydrates.

A few sulphobismuthites are found in nature:—

Kobellite. *Trisulphoplumbic sulphobismuthite* .... $\mathbf{Bi}_2Pbs''_3$.
Needle ore. *Disulphoplumbic-dicuprous sulphobismuthite* .... $\mathbf{Bi}_2Pbs''_2('Cu'_2S_2)''$.

BISMUTHOUS DITELLURO-SULPHIDE.

$Bi_2Te''_2S''$.

Sp. gr. 7·5 to 7·8.

Occurrence.—In nature, as *telluric bismuth* or *tetradymite.*

CHAPTER XVII.

MONAD ELEMENTS.

SECTION III.

POTASSIUM, K_2.

Atomic weight =39. *Probable molecular weight* =78. *Sp. gr.* 0·865. *Fuses at* 55°. *Boils at a low red heat. Atomicity* '. *Evidence of atomicity*:—

Potassic chloride	KCl.
Potassic iodide	KI.
Potassic hydrate	KHo.
Potassic sulphide	SK_2.

Occurrence.—In rocks in the form of silicate, and in soils partly as carbonate. As chloride in solid saline deposits.

In the juices of almost all plants, generally in combination with organic acids.

In sea-water and in most mineral waters.

Preparation.—1. By the action of a powerful voltaic current upon potassic hydrate, when potassium and hydrogen are liberated at the negative pole:—

$$\underset{\text{Potassic hydrate.}}{2OKH} = K_2 + H_2 + O_2.$$

2. By submitting potassic hydrate to the action of metallic iron at a strong white heat :—

$$\underset{\substack{\text{Potassic}\\\text{hydrate.}}}{4OKH} + 3Fe = \underset{\substack{\text{Magnetic iron}\\\text{oxide.}}}{{}^{iv}(Fe_3)^{viii}O_4} + 2K_2 + 2H_2.$$

3. By igniting hydric potassic tartrate (cream of tartar) out of contact with air, and subsequently mixing the residue, consisting of potassic carbonate and carbon, with charcoal, and distilling at a very high temperature :—

First operation :—

$$\underset{\substack{\text{Hydric}\\\text{potassic}\\\text{tartrate.}}}{2\left\{\begin{array}{l}COKo\\CHHo\\CHHo\\COHo\end{array}\right.} = \underset{\substack{\text{Potassic}\\\text{carbonate.}}}{COKo_2} + \underset{\text{Water.}}{5OH_2} + \underset{\substack{\text{Carbonic}\\\text{oxide.}}}{4CO} + 3C.$$

Second operation :—

$$\underset{\substack{\text{Potassic}\\\text{carbonate.}}}{COKo_2} + 2C = \underset{\substack{\text{Carbonic}\\\text{oxide.}}}{3CO} + K_2.$$

Reactions.—1. Potassium decomposes water even at its freezing-point with great energy, the heat evolved being sufficient to cause the ignition of the liberated hydrogen :—

$$K_2 + \underset{\text{Water.}}{2OH_2} = \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{2KHo} + H_2.$$

2. When potassium is ignited in a stream of carbonic anhydride, a portion of the latter is decomposed :—

$$2K_2 + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{3CO_2} = \underset{\substack{\text{Potassic}\\\text{carbonate.}}}{2COKo_2} + C.$$

COMPOUNDS OF POTASSIUM WITH CHLORINE, BROMINE, IODINE, AND FLUORINE.

Potassic chloride	KCl.
Potassic bromide	KBr.
Potassic iodide	KI.
Potassic fluoride	KF.

Potassic iodide is prepared by digesting iron filings, water, and iodine together, filtering the colourless solution, and precipitating the iron by potassic carbonate:—

$$Fe + I_2 = \underset{\text{Ferrous iodide.}}{\mathbf{Fe}I_2};$$

$$\underset{\text{Ferrous iodide.}}{\mathbf{Fe}I_2} + \underset{\text{Potassic carbonate.}}{\mathbf{C}OKo_2} = \underset{\text{Potassic iodide.}}{2KI} + \underset{\text{Ferrous carbonate.}}{\mathbf{C}OFeo''}.$$

COMPOUND OF POTASSIUM WITH HYDROXYL.

POTASSIC HYDRATE, *Caustic Potash, Potash.*

KHo, or $\mathbf{O}KH$.

Preparation.—1. By boiling in an iron vessel a solution of potassic carbonate with calcic hydrate:—

$$\underset{\text{Potassic carbonate.}}{\mathbf{C}OKo_2} + \underset{\text{Calcic hydrate.}}{\mathbf{Ca}Ho_2} = \underset{\text{Potassic hydrate.}}{2KHo} + \underset{\text{Calcic carbonate.}}{\mathbf{C}OCao''}.$$

2. By the action of potassium upon water (see p. 151).

Reactions.—By contact with acids potassic hydrate produces potassium salts:—

$$\underset{\text{Potassic hydrate.}}{KHo} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Potassic chloride.}}{KCl} + \underset{\text{Water.}}{\mathbf{O}H_2}:$$

$$\underset{\text{Potassic hydrate.}}{KHo} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Hydric potassic sulphate.}}{\mathbf{S}O_2HoKo} + \underset{\text{Water.}}{\mathbf{O}H_2}:$$

$$\underset{\text{Potassic hydrate.}}{2KHo} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Potassic sulphate.}}{\mathbf{S}O_2Ko_2} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

COMPOUNDS OF POTASSIUM WITH OXYGEN.

Potassic oxide	OK_2.	K—O—K.
Potassic dioxide	$\left\{\begin{matrix} OK \\ OK \end{matrix}\right.$	K—O—O—K.
Potassic tetroxide	$\left\{\begin{matrix} OK \\ O \\ O \\ OK \end{matrix}\right.$.	K—O—O—O—O—K.

POTASSIC OXIDE.

OK_2.

Preparation.—1. By heating potassic hydrate with potassium:—

$$\underset{\substack{\text{Potassic} \\ \text{hydrate.}}}{2KHo} + K_2 = \underset{\substack{\text{Potassic} \\ \text{oxide.}}}{2OK_2} + H_2.$$

2. By fusing together, in a current of nitrogen, potassic peroxide and potassium:—

$$\underset{\substack{\text{Potassic} \\ \text{peroxide.}}}{K_2O_4} + 3K_2 = \underset{\substack{\text{Potassic} \\ \text{oxide.}}}{4OK_2}.$$

POTASSIC DIOXIDE.

K_2O_2.

Preparation.—Obtained by the action of water on potassic peroxide.

POTASSIC TETROXIDE, *Potassic Peroxide.*

K_2O_4.

Preparation.—By fusing potassium in a current of oxygen.

COMPOUND OF POTASSIUM WITH HYDROSULPHYL.

POTASSIC SULPHHYDRATE.

KHs.

Preparation.—By saturating potassic hydrate with sulphuretted hydrogen:—

$$\underset{\text{Potassic hydrate.}}{KHo} + \underset{\text{Sulphuretted hydrogen.}}{\mathbf{S}H_2} = \underset{\text{Potassic sulphhydrate.}}{KHs} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

COMPOUNDS OF POTASSIUM WITH SULPHUR.

The following have been obtained:—

Dipotassic sulphide.........	$\mathbf{S}K_2$.	K—S—K
Dipotassic disulphide......	K_2S_2.	K—S—S—K
Dipotassic trisulphide......	K_2S_3.	K—S—S—S—K
Dipotassic tetrasulphide..	K_2S_4.	K—S—S—S—S—K.
Dipotassic pentasulphide..	K_2S_5.	K—S—S—S—S—S—K
Dipotassic heptasulphide......	K_2S_7 ?	K—S—S—S—S—S—S—S—K

DIPOTASSIC SULPHIDE.

$\mathbf{S}K_2$.

Preparation.—1. By the action of potassic hydrate on potassic sulphhydrate:—

$$\underset{\text{Potassic hydrate.}}{KHo} + \underset{\text{Potassic sulphhydrate.}}{KHs} = \underset{\text{Dipotassic sulphide.}}{\mathbf{S}K_2} + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

2. By igniting potassic sulphate with hydrogen or carbon:—

$$\underset{\text{Potassic sulphate.}}{\mathbf{S}O_2Ko_2} + 4H_2 = \underset{\text{Dipotassic sulphide.}}{\mathbf{S}K_2} + \underset{\text{Water.}}{4\mathbf{O}H_2}.$$

Reactions of dipotassic sulphide and the higher potassic sulphides:—1. By heating dipotassic sulphide with the necessary quantities of sulphur, it forms the higher potassic sulphides.

2. The potassic sulphhydrate and dipotassic sulphide, when acted upon by acids, yield sulphuretted hydrogen:—

$$\underset{\text{Potassic sulphhydrate.}}{KHs} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Potassic chloride.}}{KCl} + \underset{\text{Sulphuretted hydrogen.}}{\mathbf{S}H_2}.$$

$$\underset{\text{Dipotassic sulphide.}}{\mathbf{S}K_2} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Potassic chloride.}}{2KCl} + \underset{\text{Sulphuretted hydrogen.}}{\mathbf{S}H_2}.$$

3. The higher potassic sulphides, similarly treated, yield sulphuretted hydrogen and a precipitate of sulphur:—

$$\underset{\text{Dipotassic trisulphide.}}{K_2S_3} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Potassic chloride.}}{2KCl} + \underset{\text{Sulphuretted hydrogen.}}{\mathbf{S}H_2} + S_2.$$

4. A mixture of the higher potassic sulphides and potassic hyposulphite, known under the name of *hepar sulphuris* or *liver of sulphur*, may be prepared by heating potassic carbonate with sulphur:—

$$\underset{\text{Potassic carbonate.}}{3\mathbf{C}OKo_2} + 4S_2 = \underset{\text{Dipotassic trisulphide.}}{2K_2S_3} + \underset{\text{Potassic sulphosulphate.}}{\mathbf{S}S''OKo_2} + \underset{\text{Carbonic anhydride.}}{3\mathbf{C}O_2};$$

$$\underset{\text{Potassic carbonate.}}{3\mathbf{C}OKo_2} + 6S_2 = \underset{\text{Dipotassic pentasulphide.}}{2K_2S_5} + \underset{\text{Potassic sulphosulphate.}}{\mathbf{S}S''OKo_2} + \underset{\text{Carbonic anhydride}}{3\mathbf{C}O_2}.$$

5. The last mixture, when acted upon by acids, suffers successively the following decompositions:—

$$\underset{\text{Dipotassic pentasulphide.}}{2K_2S_5} + \underset{\text{Potassic sulphosulphate.}}{\mathbf{S}S''OKo_2} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Potassic chloride.}}{6KCl}$$

$$+ \underset{\text{Sulphosulphuric acid.}}{\mathbf{S}S''OHo_2} + \underset{\text{Sulphuretted hydrogen.}}{2\mathbf{S}H_2} + 4S_2;$$

then

$$\underset{\text{Sulphosulphuric acid.}}{\mathbf{S}S''OHo_2} = \underset{\text{Sulphurous anhydride.}}{\mathbf{S}O_2} + S + \underset{\text{Water.}}{\mathbf{O}H_2}.$$

POTASSIC CARBONATE.

$COKo_2$.

Preparation.—1. By lixiviating the ashes of land-plants.

2. By burning hydric potassic tartrate in a current of air.

SODIUM, Na_2.

Atomic weight =23. *Probable molecular weight* =46. *Sp. gr.* 0·97. *Fuses at* 90°. *Boils at a red heat.* *Atomicity* '. *Evidence of atomicity* :—

Sodic chloride	$NaCl$.
Sodic hydrate	$ONaH$.
Sodic oxide	ONa_2.

Occurrence.—In nature in the form of chloride. In sea-water and most springs. As silicate in several minerals.

Preparation.—1. By electrolyzing sodic hydrate.

2. By acting upon sodic hydrate with metallic iron at a strong white heat.

Manufacture.—By distilling in an iron retort a mixture of sodic carbonate, chalk, and charcoal :—

$$\underset{\text{Sodic carbonate.}}{CONao_2} + 2C = Na_2 + \underset{\text{Carbonic oxide.}}{3CO}.$$

Reactions.—Similar to those of potassium, but less energetic. The compounds of sodium very much resemble those of potassium.

SODIC CARBONATE.

$CONao_2$.

Manufacture.—1. Formerly by the lixiviation of the ashes of marine plants.

2. By Leblanc's process, which consists in first transforming

sodic chloride into sodic sulphate by the action of sulphuric acid:—

$$\underset{\text{Sodic chloride.}}{2NaCl} + \underset{\text{Sulphuric acid.}}{\mathbf{S}O_2Ho_2} = \underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Sodic sulphate.}}{\mathbf{S}O_2Nao_2}.$$

The sodic sulphate (technically termed *salt cake*) is next heated with calcic carbonate and small coal. The carbon reduces the sodic sulphate to sulphide; and the calcic carbonate transforms the sodic sulphide into sodic carbonate, insoluble calcic oxysulphide being simultaneously produced:—

$$\underset{\text{Sodic sulphate.}}{5\mathbf{S}O_2Nao_2} + 20C = \underset{\text{Sodic sulphide.}}{5\mathbf{S}Na_2} + \underset{\text{Carbonic oxide.}}{20\mathbf{C}O};$$

$$\underset{\text{Sodic sulphide.}}{5\mathbf{S}Na_2} + \underset{\text{Calcic carbonate.}}{7\mathbf{C}OCao''} = \underset{\text{Sodic carbonate.}}{5\mathbf{C}ONao_2}$$

$$+ \underset{\text{Calcic oxysulphide.}}{5\mathbf{C}aS'', 2\mathbf{C}aO} + \underset{\text{Carbonic anhydride.}}{2\mathbf{C}O_2}.$$

The sodic carbonate, commonly called *soda ash*, is obtained by the extraction of the resulting mass with water.

LITHIUM, Li_2.

Atomic weight =7. *Probable molecular weight* =14. *Sp. gr.* =0·59. *Fuses at* 180°. *Atomicity* $'$. *Evidence of atomicity*:—

Lithic chloride .... $LiCl$.
Lithic hydrate (Lithia) .... $\mathbf{O}LiH$.

Occurrence.—In nature, in the minerals *petalite, spodumene, lepidolite,* and *triphylline*, and in small quantities in some mineral waters and ashes of plants.

The properties of lithium resemble those of potassium and sodium; and the compounds of the three metals also exhibit considerable similarity.

SECTION IV.

SILVER, Ag_2.

Atomic weight $=108$. *Probable molecular weight* $=216$. *Sp. gr.* 10·4743. *Fuses at about* 1000°. *Atomicity* '. *Evidence of atomicity*:—

Argentic chloride	$AgCl$.
Argentic iodide	AgI.
Argentic oxide	OAg_2.

Occurrence.—In nature in the free state, and as sulphide in *silver glance*; as sulphantimonite in *dark-red silver-ore* (see p. 142), as chloride in *horn-silver*, as a compound of bromide and chloride ($2AgBr, 3AgCl$) in *embolite*, and also as carbonate.

Extraction.—1. The silver minerals are roasted with sodic chloride, by which the metal is converted into chloride; the mass is then mixed with water, scrap iron, and mercury, and agitated for some hours. The iron reduces the argentic chloride to the metallic state, and the silver is then dissolved by the mercury.

2. By crystallizing argentiferous lead. Nearly pure lead is first deposited; and the residue, rich in silver, is then cupelled.

Reactions.—1. Silver is blackened by sulphuretted hydrogen, argentic sulphide being formed.

2. Silver is acted upon by hot concentrated sulphuric acid:—

$$Ag_2 + \underset{\text{Sulphuric acid.}}{2SO_2Ho_2} = \underset{\text{Argentic sulphate.}}{SO_2Ago_2} + \underset{\text{Water.}}{2OH_2} + \underset{\text{Sulphurous anhydride.}}{SO_2}.$$

3. Dilute nitric acid readily dissolves silver:—

$$3Ag_2 + \underset{\text{Nitric acid.}}{8NO_2Ho} = \underset{\text{Argentic nitrate.}}{6NO_2Ago} + \underset{\text{Water.}}{4OH_2} + \underset{\text{Nitric oxide.}}{N''_2O_2}.$$

4. At a red heat silver decomposes hydrochloric acid:—

$$Ag_2 + \underset{\text{Hydrochloric acid.}}{2HCl} = H_2 + \underset{\text{Argentic chloride.}}{2AgCl}.$$

There are three compounds of silver with oxygen:—

Argentous oxide	OAg_4?
Argentic oxide	OAg_2.
Argentic peroxide	$\left\{\begin{matrix}OAg\\OAg\end{matrix}\right.$.

Argentous oxide is prepared by heating argentic citrate to 100° in a stream of hydrogen, dissolving the residue, which contains argentous citrate, in cold water, and precipitating the argentous oxide by potassic hydrate.

Argentic oxide is formed by precipitating argentic nitrate with a solution of baric hydrate, and drying the precipitate, which is probably argentic hydrate, AgHo. This is the salifiable oxide of silver.

Argentic peroxide is obtained by electrolyzing a solution of argentic nitrate, when it is deposited upon the positive pole.

Argentic chloride, bromide, and iodide are insoluble in water and nitric acid.

CHAPTER XVIII.

DYAD ELEMENTS.

SECTION II.

BARIUM, Ba.

Atomic weight = 137. *Probable molecular weight* = 137. *Sp. gr. between* 4·0 *and* 5·0. *Fuses below a red heat. Atomicity''. Evidence of atomicity:—*

Baric chloride	$Ba''Cl_2$.
Baric hydrate	$Ba''Ho_2$.
Baric oxide	$Ba''O$.

Occurrence.—In nature in the form of sulphate in the mineral *heavy spar*, and as carbonate ($\mathbf{C}OBao''$) in *witherite.*

Preparation.—1. By electrolyzing moistened baric hydrate, carbonate, nitrate, or chloride, the negative electrode being mercury. An amalgam of barium is thus formed, from which the mercury is removed by distillation.

2. By passing the vapour of potassium or sodium over baric oxide strongly heated in an iron tube, and extracting the metal by means of mercury.

3. By acting upon a solution of baric chloride with sodium amalgam, barium amalgam is produced.

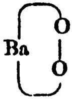
Reaction.—Barium decomposes water at the common temperature :—

$$\mathrm{Ba} + \underset{\text{Water.}}{2\mathbf{O}H_2} = H_2 + \underset{\text{Baric hydrate.}}{\mathbf{Ba}Ho_2}.$$

COMPOUNDS OF BARIUM WITH OXYGEN.

Baric oxide .... $\mathbf{Ba}O$. $Ba{=}O$

Baric peroxide .... $\mathbf{Ba}\begin{matrix}O\\O\end{matrix}\Big\}$.



BARYTA, *Baric Oxide.*

$\mathbf{Ba}O$.

Preparation.—1. By converting the native carbonate into nitrate by the action of nitric acid, and then heating the nitrate to redness in an iron crucible :—

$$\underset{\text{Baric carbonate.}}{\mathbf{C}\mathrm{OBao}''} + \underset{\text{Nitric acid.}}{2\mathbf{N}\mathrm{O_2Ho}} = \underset{\text{Baric nitrate.}}{\left\{\begin{matrix}\mathbf{N}\mathrm{O_2}\\ \mathrm{Bao}''\\ \mathbf{N}\mathrm{O_2}\end{matrix}\right.} + \underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}} + \underset{\text{Carbonic anhydride.}}{\mathbf{C}\mathrm{O_2}};$$

$$\underset{\text{Baric nitrate.}}{2\left\{\begin{matrix}\mathbf{N}\mathrm{O_2}\\ \mathrm{Bao}''\\ \mathbf{N}\mathrm{O_2}\end{matrix}\right.} = \underset{\text{Baric oxide.}}{2\mathbf{Ba}\mathrm{O}} + \underset{\text{Nitric peroxide.}}{2\mathbf{N}^{iv}_2\mathrm{O_4}} + \mathrm{O_2}.$$

2. The nitrate may be obtained from native baric sulphate by mixing the latter with charcoal and heating the mixture to a high red heat, by which the sulphate is converted into sulphide :—

$$\underset{\text{Baric sulphate.}}{\mathbf{S}\mathrm{O_2Bao}''} + \mathrm{C_4} = \underset{\text{Baric sulphide.}}{\mathbf{Ba}\mathrm{S}''} + \underset{\text{Carbonic oxide.}}{4\mathbf{C}\mathrm{O}};$$

the residue is then treated with dilute nitric acid, when baric nitrate is formed.

Reaction.—In contact with water, baric oxide is converted, with great evolution of heat, into baric hydrate :—

$$\underset{\text{Baric oxide.}}{\mathbf{Ba}\mathrm{O}} + \underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}} = \underset{\text{Baric hydrate.}}{\mathbf{Ba}\mathrm{Ho_2}}.$$

BARIC PEROXIDE.

$$\mathbf{Ba}\left.\begin{matrix}\mathrm{O}\\ \mathrm{O}\end{matrix}\right\}.$$

Preparation.—1. By passing oxygen over baric oxide or baric hydrate heated to dull redness :—

$$\underset{\text{Baric oxide.}}{2\mathbf{Ba}\mathrm{O}} + \mathrm{O_2} = \underset{\text{Baric peroxide.}}{2\mathbf{Ba}\left.\begin{matrix}\mathrm{O}\\ \mathrm{O}\end{matrix}\right\}}.$$

$$\underset{\text{Baric hydrate.}}{2\mathbf{Ba}\mathrm{Ho_2}} + \mathrm{O_2} = \underset{\text{Baric peroxide.}}{2\mathbf{Ba}\left.\begin{matrix}\mathrm{O}\\ \mathrm{O}\end{matrix}\right\}} + \underset{\text{Water.}}{2\mathbf{O}\mathrm{H_2}}.$$

2. By heating baric oxide to redness in a crucible and gradually adding potassic chlorate:—

$$\underset{\substack{\text{Baric}\\\text{oxide.}}}{3\mathbf{Ba}O} + \underset{\substack{\text{Potassic}\\\text{chlorate.}}}{ClO_2Ko} = \underset{\substack{\text{Potassic}\\\text{chloride.}}}{KCl} + \underset{\substack{\text{Baric}\\\text{peroxide.}}}{3\mathbf{Ba}{\begin{matrix}O\\O\end{matrix}}\Big\}}.$$

Reactions.—1. By the action of heat it splits into baric oxide and oxygen:—

$$\underset{\substack{\text{Baric}\\\text{peroxide.}}}{2\mathbf{Ba}{\begin{matrix}O\\O\end{matrix}}\Big\}} = \underset{\substack{\text{Baric}\\\text{oxide.}}}{2\mathbf{Ba}O} + O_2.$$

2. By treatment with steam, at the same temperature at which the peroxide was previously formed, it produces baric hydrate and oxygen:—

$$\underset{\substack{\text{Baric}\\\text{peroxide.}}}{2\mathbf{Ba}{\begin{matrix}O\\O\end{matrix}}\Big\}} + \underset{\text{Water.}}{2\mathbf{O}H_2} = \underset{\substack{\text{Baric}\\\text{hydrate.}}}{2\mathbf{Ba}Ho_2} + O_2.$$

3. By the action of carbonic anhydride or acids upon baric peroxide, hydroxyl is formed (p. 45).

COMPOUND OF BARIUM WITH HYDROXYL.

BARIC HYDRATE, *Caustic Baryta.*

$\mathbf{Ba}Ho_2$.

Preparation.—1. By the action of water on baric oxide (p. 161).

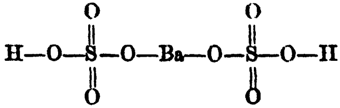
2. By boiling in water, with cupric oxide, the mass containing baric sulphide, prepared by reducing baric sulphate with carbon:—

$$\underset{\substack{\text{Baric}\\\text{sulphide.}}}{\mathbf{Ba}S''} + \underset{\substack{\text{Cupric}\\\text{oxide.}}}{\mathbf{Cu}O} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\substack{\text{Baric}\\\text{hydrate.}}}{\mathbf{Ba}Ho_2} + \underset{\substack{\text{Cupric}\\\text{sulphide.}}}{\mathbf{Cu}S''}.$$

Barium salts are formed by the action of acids upon baric hydrate, carbonate, or oxide.

DIHYDRIC BARIC DISULPHATE.

$$\left\{\begin{array}{l}\mathbf{S}O_2Ho\\ Bao''\\ \mathbf{S}O_2Ho\end{array}\right. .$$



$$H-O-\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}-O-Ba-O-\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}-O-H$$

This compound is formed by boiling baric sulphate in concentrated sulphuric acid, when the salt crystallizes on cooling.

STRONTIUM, Sr.

Atomic weight =87·5. *Probable molecular weight* =87·5. *Sp. gr.* 2·5. *Fuses at a higher temperature than barium.* *Atomicity* ''. *Evidence of atomicity* :—

Strontic chloride	$\mathbf{Sr}''Cl_2$.
Strontic hydrate	$\mathbf{Sr}''Ho_2$.
Strontic oxide	$\mathbf{Sr}''O$.

Occurrence.—In the form of carbonate as the mineral *strontianite*, and as sulphate in *celestine.*

Preparation.—1. By the same methods as those employed in the preparation of barium.

2. By electrolyzing fused strontic chloride.

Character.—The compounds of strontium resemble those of barium in constitution, preparation, and properties.

The ***strontic peroxide*** can only be prepared by adding hydroxyl to a solution of strontic hydrate :—

$$\underset{\text{Strontic hydrate.}}{\mathbf{Sr}Ho_2} + \underset{\text{Hydroxyl.}}{Ho_2} = \underset{\text{Strontic peroxide.}}{\mathbf{Sr}\left.\begin{matrix}O\\O\end{matrix}\right\}} + \underset{\text{Water.}}{2\mathbf{O}H_2}.$$

Strontic carbonate is more easily decomposed by heat than baric carbonate.

CALCIUM, Ca.

Atomic weight = 40. *Probable molecular weight* = 40. *Sp. gr.* 1·6. *Atomicity''*. *Evidence of atomicity*:—

Calcic chloride	$\mathbf{Ca}''\mathrm{Cl}_2$.
Calcic hydrate	$\mathbf{Ca}''\mathrm{Ho}_2$.
Calcic oxide	$\mathbf{Ca}''\mathrm{O}$.

Occurrence.—In nature as carbonate in the numerous *calc spars*, chalk, marble, &c.; as tetrahydric calcic sulphate ($\mathbf{S}\mathrm{Ho}_4\mathrm{Cao}''$) in *gypsum*, *alabaster*, *selenite*, &c.; as phosphate in *apatite* and *phosphorite* (see p. 123); as fluoride in the *fluor spars* (see p. 98); and in combination with silicon, oxygen, and other metals in numerous minerals.

Preparation.—1. By processes similar to those employed for the preparation of barium and strontium.

2. By fusing together sodium, zinc, and calcic chloride, and subsequently heating the alloy of calcium and zinc so obtained, to a very high temperature in a crucible of gas-carbon, when the zinc volatilizes, leaving the calcium, which contains, however, a small quantity of iron.

Character.—The compounds of calcium resemble those of barium and strontium.

Calcic oxide or *quicklime* ($\mathbf{Ca}\mathrm{O}$) is manufactured on a large scale by burning coal intermixed with chalk or limestone, when carbonic anhydride is easily expelled from the chalk or limestone, leaving calcic oxide.

Calcic hydrate or *slaked lime* ($\mathbf{Ca}\mathrm{Ho}_2$) is formed by the action of water upon calcic oxide; it is much less soluble in water than the baric and strontic hydrates.

By exposing quicklime to the action of the air it is converted into *dicalcic carbonate dihydrate*, $\mathbf{C}\mathrm{O}(\mathrm{OCa}''\mathrm{Ho})_2$.

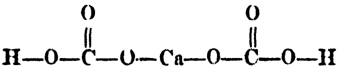
Calcic peroxide $\left(\mathbf{Ca}\begin{matrix}\mathrm{O}\\\mathrm{O}\end{matrix}\right\}\Big)$ is prepared like the corresponding strontium compound.

By passing chlorine over calcic hydrate, a compound known as ***chloride of lime*** or ***bleaching-powder*** is formed. This has been supposed to consist of calcic chloride mixed with calcic hypochlorite; but it is more probably calcic chloro-hypochlorite, as expressed by the following formulæ:—

$$\mathbf{Ca}(OCl)Cl. \qquad Cl—Ca—O—Cl.$$

The corresponding baric and strontic chloro-hypochlorites $\mathbf{Ba}\begin{matrix}(OCl)\\Cl\end{matrix}$ and $\mathbf{Sr}\begin{matrix}(OCl)\\Cl\end{matrix}$, are known.

Barium, strontium, and calcium all form soluble dihydric dicarbonates:—



$$H—O—\overset{\overset{O}{\|}}{C}—O—Ca—O—\overset{\overset{O}{\|}}{C}—O—H$$

They are produced by passing an excess of carbonic anhydride through solutions of baric, strontic, and calcic hydrates. The compounds are decomposed at 100°, carbonic anhydride being evolved and carbonates precipitated:—

$$\underset{\text{Dihydric calcic dicarbonate.}}{\left\{\begin{matrix}\mathbf{C}OHo\\Cao''\\\mathbf{C}OHo\end{matrix}\right.} = \underset{\text{Calcic carbonate.}}{\mathbf{C}OCao''} + \underset{\text{Water.}}{\mathbf{O}H_2} + \underset{\text{Carbonic anhydride.}}{\mathbf{C}O_2}.$$

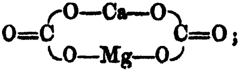
MAGNESIUM, Mg.

Atomic weight =24. *Probable molecular weight* =24. *Sp. gr.* 1·75. ***Fuses at a red heat. Volatilizes at a bright red heat. Atomicity* ''. *Evidence of atomicity*:—**

Magnesic chloride	$\mathbf{Mg}''Cl_2$.
Magnesic oxide	$\mathbf{Mg}''O$.
Magnesic hydrate	$\mathbf{Mg}''Ho_2$.

Occurrence.—In nature in *dolomite,* the calcic magnesic dicarbonate,

$\begin{matrix}\mathbf{C}O\\ \mathbf{C}O\end{matrix}Cao''Mgo''$



$O{=}C\begin{matrix}\diagup O-Ca-O\diagdown\\ \diagdown O-Mg-O\diagup\end{matrix}C{=}O$;

in *brucite* or magnesic hydrate, $\mathbf{Mg}Ho_2$, and in many minerals containing silicon.

Preparation.—1. By electrolyzing fused magnesic chloride.

2. By fusing a mixture of magnesic chloride, potassic chloride, and sodium.

Reactions.—1. It very slowly decomposes water at the ordinary temperature, but more rapidly at a boiling heat.

2. It readily burns when heated to redness in the air.

Character.—Magnesium only forms one compound with oxygen, $\mathbf{Mg}O$, *magnesia*. It is obtained by burning magnesium in air, or by heating the carbonate to redness.

Magnesic hydrate ($\mathbf{Mg}Ho_2$) is formed by the action of water upon magnesic oxide, or by precipitating magnesic sulphate by potassic hydrate:—

$$\underset{\substack{\text{Magnesic}\\ \text{sulphate.}}}{\mathbf{S}O_2Mgo''} + \underset{\substack{\text{Potassic}\\ \text{hydrate.}}}{2\mathbf{O}KH} = \underset{\substack{\text{Potassic}\\ \text{sulphate.}}}{\mathbf{S}O_2Ko_2} + \underset{\substack{\text{Magnesic}\\ \text{hydrate.}}}{\mathbf{Mg}Ho_2}.$$

It scarcely dissolves in water.

Crystallized magnesic sulphate ($\mathbf{S}OHo_2Mgo''$, $6\mathbf{O}H_2$) is prepared by treating *dolomite*, the magnesic calcic dicarbonate, with sulphuric acid, filtering from the nearly insoluble calcic sulphate, and crystallizing:—

$$\underset{\substack{\text{Dolomite, calcic}\\ \text{magnesic dicarbonate.}}}{\begin{matrix}\mathbf{C}O\\ \mathbf{C}O\end{matrix}Cao''Mgo''} + \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{2\mathbf{S}O_2Ho_2} = \underset{\substack{\text{Dihydric calcic}\\ \text{sulphate.}}}{\mathbf{S}OHo_2Cao''} + \underset{\substack{\text{Dihydric magnesic.}\\ \text{sulphate.}}}{\mathbf{S}OHo_2Mgo''}$$

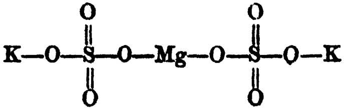
$$+ \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{2\mathbf{C}O_2}.$$

Magnesic sulphate is very soluble in water, thus differing from the baric, strontic, and calcic sulphates.

Magnesic sulphate, when mixed with potassic or ammonic sulphate, forms a disulphate, as, for instance,

$$\left.\begin{array}{l}SO_2Ko\\SO_2Ko\end{array}\right\}Mgo''.$$

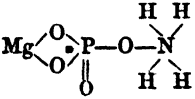
Dipotassic magnesic disulphate.



Many magnesic phosphates are known.

Ammonic magnesic phosphate,

$POAmoMgo''$,

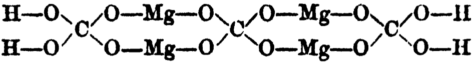


occurs in the seeds of some of the cereals, and sometimes in urine, and in urinary calculi: it is found in nature as *guanite* and *struvite*.

Magnesic carbonate ($COMgo''$) is found in nature as *magnesite.*

MAGNESIA ALBA, *Tetrahydric tetramagnesic tricarbonate.*

$$C_3Mgo''_4Ho_4, \text{ or} \left\{\begin{array}{l}CHo_2\\Mgo''_2\\C\\Mgo''_2\\CHo_2\end{array}\right.,$$



This compound is formed by boiling a solution of magnesic sulphate with sodic carbonate (Berzelius):—

$$\underset{\text{Magnesic sulphate.}}{4SO_2Mgo''} + \underset{\text{Sodic carbonate.}}{4CONao_2} + \underset{\text{Water.}}{2OH_2} = \underset{\text{Magnesia alba.}}{C_3Mgo''_4Ho_4}$$

$$+ \underset{\text{Sodic sulphate.}}{4SO_2Nao_2} + \underset{\text{Carbonic anhydride.}}{CO_2}$$

ZINC, Zn.

Atomic weight = 65. *Molecular weight* = 65. *Molecular and atomic volume* ☐☐. 1 *litre of zinc vapour weighs* 32·5 *criths. Sp. gr.* 6·8 *to* 7·2. *Fuses at* 500°. *Distils at a red heat. Atomicity* ''. *Evidence of atomicity*:—

Zincic chloride	$\mathbf{Zn}''Cl_2$.
Zincic oxide	$\mathbf{Zn}''O$.
Zincic hydrate	$\mathbf{Zn}''Ho_2$.

Occurrence.—In nature as oxide ($\mathbf{Zn}O$) in *red zinc*, as sulphide ($\mathbf{Zn}S''$) in the mineral *zinc blende*, as carbonate ($\mathbf{C}OZno''$) in *calamine*, and as silicate in *electric calamine*, *williamite*, or *zinc glass*.

Manufacture.—Zinc blende, or calamine, by being roasted in a current of air, is converted into zincic oxide:—

$$\underset{\text{Zincic carbonate (calamine).}}{\mathbf{C}OZno''} = \underset{\text{Zincic oxide.}}{\mathbf{Zn}O} + \underset{\text{Carbonic anhydride.}}{\mathbf{C}O_2}.$$

$$\underset{\text{Zincic sulphide (Zinc blende).}}{2\mathbf{Zn}S''} + 3O_2 = \underset{\text{Zincic oxide.}}{2\mathbf{Zn}O} + \underset{\text{Sulphurous anhydride.}}{2\mathbf{S}O_2}.$$

The roasted and powdered mineral is then heated with powdered coal, when the zinc is reduced and distils over:—

$$\underset{\text{Zincic oxide.}}{\mathbf{Zn}O} + C = Zn + \underset{\text{Carbonic oxide.}}{\mathbf{C}O}.$$

Reactions.—1. It slowly decomposes aqueous vapour at 100°:—

$$\underset{\text{Water.}}{\mathbf{O}H_2} + Zn = \underset{\text{Zincic oxide.}}{\mathbf{Zn}O} + H_2.$$

2. Zinc is attacked by almost every acid at the common temperature.

3. When boiled in potassic, sodic, or even ammonic hydrate, hydrogen is evolved, and a double oxide formed:—

$$\underset{\text{Potassic hydrate.}}{2OKH} + Zn = \underset{\text{Dipotassic zincic oxide.}}{ZnKo_2} + H_2.$$

COMPOUND OF ZINC WITH OXYGEN.

ZINCIC OXIDE.

ZnO.

Preparation.—1. Zincic oxide is obtained by burning zinc in air.

2. By passing steam over heated zinc.

3. By heating the precipitate formed by ammonic carbonate in solutions of zinc salts.

OTHER COMPOUNDS OF ZINC.

Zincic hydrate ($ZnHo_2$) is obtained as a white precipitate by the action of potassic hydrate on solutions of zinc salts:—

$$\underset{\text{Zincic sulphate.}}{SO_2Zno''} + \underset{\text{Potassic hydrate.}}{2OKH} = \underset{\text{Zincic hydrate.}}{ZnHo_2} + \underset{\text{Potassic sulphate.}}{SO_2Ko_2}.$$

The precipitate is dissolved by excess of potassic hydrate.

Crystallized zincic sulphate is isomorphous with crystallized magnesic sulphate, and contains seven molecules of water, six of which are easily expelled at a moderate heat, the last only being driven off at a somewhat high temperature. It also resembles magnesic sulphate in forming double salts with potassic and ammonic sulphates:—

Zincic sulphate (crystallized) .... $SOHo_2Zno'', 6OH_2$.

Dipotassic zincic disulphate (crystallized) $\left\{\begin{array}{l} SO_2Ko \\ Zno'' \\ SO_2Ko \end{array}\right., 6OH_2.$

Zincic carbonate **($\mathbf{C}OZno''$) occurs in nature as *calamine*.**

The precipitate obtained by adding a solution of sodic carbonate to a solution of a salt of zinc has a variable constitution. The reaction usually takes place thus:—

$$\underset{\text{Zincic sulphate.}}{5\mathbf{S}O_2Zno''} + \underset{\text{Sodic carbonate.}}{5\mathbf{C}ONao_2} + \underset{\text{Water.}}{3\mathbf{O}H_2} = \underset{\substack{\text{Dihydric pentazincic}\\ \text{dicarbonate tetrahydrate.}}}{\left\{\begin{array}{l}\mathbf{C}Ho(OZn''Ho)_2\\ Zno''\\ \mathbf{C}Ho(OZn''Ho)_2\end{array}\right.}$$

$$+ \underset{\text{Sodic sulphate.}}{5\mathbf{S}O_2Nao_2} + \underset{\text{Carbonic anhydride.}}{3\mathbf{C}O_2}.$$

CHAPTER XIX.

DYAD ELEMENTS.

SECTION IV.

CADMIUM, Cd.

Atomic weight =112. *Molecular weight* =112. *Molecular and atomic volume* ☐☐. 1 *litre of cadmium vapour weighs* 56 *criths*. *Sp. gr.* 8·7. *Fuses below* 260°. *Easily volatile*. *Atomicity* ''. *Evidence of atomicity*:—

Cadmic chloride.................... $\mathbf{Cd}''Cl_2$.
Cadmic oxide .... $\mathbf{Cd}''O$.

***Occurrence*.—In nature in small quantities, associated with zinc, and in the form of sulphide as *greenockite*.**

***Preparation*.—By distilling fractionally the more volatile part of the metal obtained in the manufacture of zinc, and then dissolving this more volatile product (which consists of**

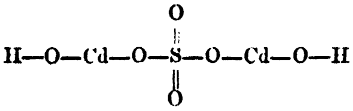
zinc, cadmium, and a little copper) in hydrochloric or dilute sulphuric acid, precipitating the cadmium and copper with sulphuretted hydrogen, dissolving the mixed sulphides in dilute sulphuric acid, and adding an excess of solution of ammonic carbonate, which precipitates both cadmium and copper, but redissolves the latter. The cadmic carbonate is then ignited, and the resulting oxide reduced by charcoal.

Cadmic oxide ($\mathbf{Cd}O$) is prepared by heating the hydrate, carbonate, or nitrate.

Cadmic hydrate ($\mathbf{Cd}Ho_2$) is obtained by precipitating a solution of a cadmic salt by sodic or potassic hydrate.

Cadmic sulphate ($\mathbf{S}O_2Cdo'', 4\mathbf{O}H_2$) is obtained by dissolving cadmic oxide or carbonate in sulphuric acid. By heating this compound, or by partially decomposing it with alkaline hydrates, it is transformed into

Dicadmic sulphate dihydrate .... $\mathbf{S}O_2(OCd''Ho)_2$



$$\begin{matrix} & & & & & & O & & & & & & \\ & & & & & & \| & & & & & & \\ H & — & O & — & Cd & — & O—S & — & O & — & Cd & — & O—H \\ & & & & & & \| & & & & & & \\ & & & & & & O & & & & & & \end{matrix}$$

MERCURY, Hg.

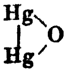
Atomic weight =200. *Molecular weight* =200. *Molecular and atomic volume* ☐☐. 1 *litre of mercury vapour weighs* 100 *criths*. *Sp. gr.* 13·59. *Fuses at* −40°. *Boils at* 360°. *Atomicity* ″, *also a pseudo-monad.*

The following list contains the names and probable constitutional formulæ of the principal compounds of this metal:—

Mercurous chloride (*horn-mercury*) } $\mathbf{Hg}'_2Cl_2$, or $\left\{\begin{matrix}\mathbf{Hg}Cl\\ \mathbf{Hg}Cl\end{matrix}\right.$. Cl—Hg—Hg—Cl

Mercuric chloride... $\mathbf{Hg}Cl_2$.

Mercurous oxide... $\mathbf{Hg}'_2O$, or $\left\{\begin{matrix}\mathbf{Hg}\\ \mathbf{Hg}\end{matrix}\right\}O$.

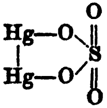


Mercuric oxide ... $\mathbf{Hg}O$. $Hg{=}O$

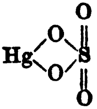
Mercurous sulphide $\mathbf{Hg}'_2S''$, or $\left\{\begin{matrix}\mathbf{Hg}\\ \mathbf{Hg}\end{matrix}\right.S''$.

Mercuric sulphide (*vermilion, cinnabar*) .... $\mathbf{Hg}S''$.

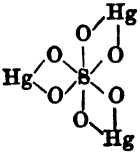
Mercurous sulphate. $\mathbf{S}O_2('Hg'_2O_2)''$.



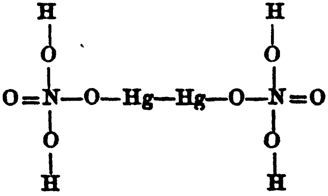
Mercuric sulphate... $\mathbf{S}O_2Hgo''$.



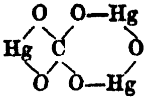
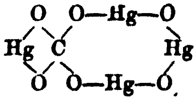
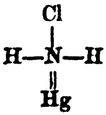
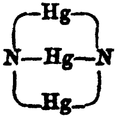
Trimercuric sulphate (*Turpeth mineral*). $\mathbf{S}Hgo''_3$.



Tetrahydric mercurous dinitrate... $\mathbf{N}_2O_2Ho_4('Hg'_2O_2)''$.



Dimercurous dinitrate ... $\left\{\begin{matrix}\mathbf{N}O('Hg'_2O_2)'\\ O\\ \mathbf{N}O('Hg'_2O_2)''\end{matrix}\right.$.

Hexahydric trimercurous tetranitrate …	$\mathbf{N}_4O_4Ho_6('Hg'_2O_2)''_3$.	
Mercurous dimercuric dinitrate …………	$\mathbf{N}_2O_2('Hg'_2O_2)''Hgo''_2$.	
Tetrahydric mercuric dinitrate …………	$\mathbf{N}_2O_2Ho_4Hgo''$.	
Tetrahydric dimercuric dinitrate …………	$\mathbf{N}_2OHo_4Hgo''_2$.	
Dihydric trimercuric dinitrate …………	$\mathbf{N}_2OHo_2Hgo''_3$.	
Trimercuric carbonate…	$\mathbf{C}Hgo''(Hg''_2O_2)''$.	
Tetramercuric carbonate.	$\mathbf{C}Hgo''(Hg''_3O_4)''$.	
Mercurosodiammonic dichloride …………	$\left.\begin{matrix}\mathbf{N}H_3ClHg\\ \mathbf{N}H_3ClHg\end{matrix}\right\}$.	
Mercurosomercurodiammonic dichloride.	$\mathbf{N}_2Hg''H_4'Hg'Cl_2$.	
Mercurammonic chloride. (*White precipitate.*) ……………	$\mathbf{N}H_2Hg''Cl$.	
Trimercuric diamide …	$\mathbf{N}_2Hg''_3$.	

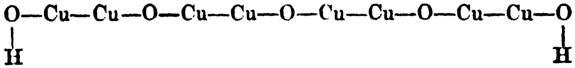
COPPER, Cu.

Atomic weight =63·5. *Probable molecular weight* =63·5. *Sp. gr.* 8·8. *Fuses at about* 780°. *Atomicity* '', *also a pseudo-monad.*

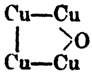
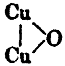
The following are the names and probable constitutional formulæ of the principal compounds of this metal :—

Cuprous hydride.........	$\left\{\begin{matrix}\mathbf{Cu}H\\ \mathbf{Cu}H\end{matrix}\right.$	H—Cu—Cu—H
Cuprous chloride ...	$'\mathbf{Cu}'_2Cl_2$ or $\left\{\begin{matrix}\mathbf{Cu}Cl\\ \mathbf{Cu}Cl\end{matrix}\right.$	Cl—Cu—Cu—Cl.
Cupric chloride	$\mathbf{Cu}Cl_2, 2\mathbf{O}H_2$.	Cl—Cu—Cl

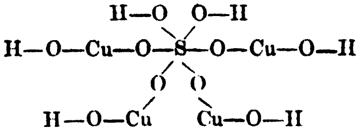
Cuprous hydrate, $4'\mathbf{Cu}'_2O, \mathbf{O}H_2$, or $\left\{\begin{array}{l}\mathbf{Cu}Ho\\ \left.\begin{array}{l}\mathbf{Cu}\\ \mathbf{Cu}\end{array}\right\}O\\ \left.\begin{array}{l}\mathbf{Cu}\\ \mathbf{Cu}\end{array}\right\}O.\\ \left.\begin{array}{l}\mathbf{Cu}\\ \mathbf{Cu}\end{array}\right\}O\\ \mathbf{Cu}Ho\end{array}\right.$



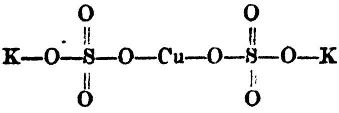
$$\begin{array}{l}O—Cu—Cu—O—Cu—Cu—O—Cu—Cu—O—Cu—Cu—O\\ \,|\hspace{27em}|\\ H\hspace{27em}H\end{array}$$

Cupric hydrate	$\mathbf{Cu}Ho_2$.	H—O—Cu—O—H
Cuprous quadrantoxide.	$\left\{\begin{array}{l}\mathbf{Cu}\\ Cu''\\ Cu''\\ \mathbf{Cu}\end{array}\right]O$	
Cuprous oxide. (*Red copper ore*, or *ruby ore*.)	$'\mathbf{Cu}'_2O$ or $\left\{\begin{matrix}\mathbf{Cu}\\ \mathbf{Cu}\end{matrix}\right\}O$.	
Cupric oxide	$\mathbf{Cu}O$.	Cu=O
Cuprous sulphide. (*Copper glance.*) ...	$'\mathbf{Cu}'_2S''$ or $\left\{\begin{matrix}\mathbf{Cu}\\ \mathbf{Cu}\end{matrix}\right\}S''$.	

Cupric sulphide. (*Indigo copper* or *blue copper.*)	$\mathbf{Cu}S''$.
Cupric sulpho-hydrate..	$5\mathbf{Cu}S'', \mathbf{Cu}Ho_2$.
Cupric nitrate	$\begin{matrix}\mathbf{N}O_2\\ \mathbf{N}O_2\end{matrix}Cuo'', 4\mathbf{O}H_2$, or $\begin{matrix}\mathbf{N}Ho_4\\ \mathbf{N}Ho_4\end{matrix}Cuo''$.
Dihydric cupric sulphate	$\mathbf{S}OHo_2Cuo'', 4\mathbf{O}H_2$.
Hydric tricupric sulphate trihydrate...	$\mathbf{S}OHo(OCu''Ho)_3$.
Dihydric tetracupric sulphate tetrahydrate. (*Brochantite.*)	$\mathbf{S}Ho_2(OCu''Ho)_4$.



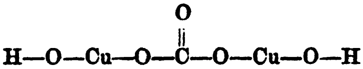
Hydric pentacupric sulphate pentahydrate	$\mathbf{S}Ho(OCu''Ho)_5, 2\mathbf{O}H_2$.
Ammoniocupric sulphate. (*Dihydric diammonic cuprodiammonic sulphate.*)	$\mathbf{S}Ho_2Amo_2 \left[\begin{matrix}N^vH_3\\ N^vH_3\end{matrix}CuO''_2\right]''$
Dipotassic cupric disulphate	$\left\{\begin{matrix}\mathbf{S}O_2Ko\\ Cuo''\\ \mathbf{S}O_2Ko\end{matrix}\right.$.



$$K-O-\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}-O-Cu-O-\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}-O-K$$

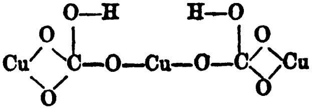
Dicupric carbonate. (*Mysorin.*)	$\mathbf{C}Cuo''_2$.

Dicupric carbonate dihydrate. (*Malachite.*) $\mathrm{CO(OCu''Ho)_2}$.

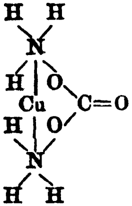


$$\mathrm{H{-}O{-}Cu{-}O{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}O{-}Cu{-}O{-}H}$$

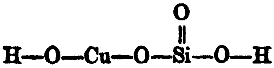
Dihydric tricupric dicarbonate. (*Blue malachite, azurite, mountain-blue,* or *copper-azure.*) …… $\left\{\begin{array}{l}\mathrm{CHoCuo''}\\ \mathrm{Cuo''}\\ \mathrm{CHoCuo''}\end{array}\right.$.



Cuprodiammonic carbonate. (*Ammoniocupric carbonate.*) ………… $\mathrm{CO}\left[\begin{array}{l}\mathrm{N^vH_3}\\ \mathrm{N^vH_3}\end{array}\mathrm{Cu''O_2}\right]''$.



Hydric cupric silicate hydrate. (*Dioptase.*) $\mathrm{SiOHo(OCu''Ho)}$.



$$\mathrm{H{-}O{-}Cu{-}O{-}\overset{\overset{\displaystyle O}{\|}}{Si}{-}O{-}H}$$

CHAPTER XX.

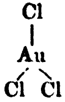
TRIAD ELEMENTS.

SECTION II.

GOLD, Au_2.

Atomic weight=196·7. *Probable molecular weight*=393·4. *Sp. gr.* 19·3 *to* 19·5. *Fuses at about* 1100°-1200°. *Atomicity*' *and* '''.

The following are the names and probable constitutional formulæ of the chief compounds of gold :—

Aurous chloride	$AuCl$.	$Au—Cl$.
Auric chloride	$AuCl_3$.	 $\begin{matrix} & Cl & \\ & \vert & \\ & Au & \\ Cl & & Cl \end{matrix}$
Aurous iodide	AuI.	
Auric iodide	AuI_3.	
Aurous oxide	Au_2O.	$Au—O—Au$
Auric oxide. (*Auric anhydride.*)	$\left\{\begin{matrix}AuO\\O\\AuO\end{matrix}\right.$.	$O{=}Au—O—Au{=}O$
Potassic aurate	$AuOKo, 3OH_2$.	$O{=}Au—O—K$
Aurous sulphide...........	Au_2S''.	
Auric sulphide	$\left\{\begin{matrix}AuS''\\S''\\AuS''\end{matrix}\right.$.	

SECTION III.

THALLIUM, Tl_2.

Atomic weight =204. *Probable molecular weight* =408. *Sp. gr.* 11·81 to 11·91. *Fuses at* 294°. *Atomicity* $'$ and $'''$. *Evidence of atomicity*:—

Thallous chloride	$TlCl$.
Thallous oxide	OTl_2.
Thallic chloride	$Tl'''Cl_3$.

Occurrence.—In small quantities in certain varieties of pyrites and sulphur, and in minute quantities in some mineral springs.

Preparation.—By extracting with water the deposit formed in the flues of sulphuric acid-chambers, and precipitating the thallous chloride by hydrochloric acid. The chloride is converted into sulphate by the action of sulphuric acid; and when purified, a solution of the sulphate is decomposed by metallic zinc, which precipitates the metallic thallium.

The following list contains the principal compounds of this metal:—

Thallous chloride	$TlCl$.
Thallous oxide	OTl_2.
Thallic chloride	$TlCl_3$.
Thallic oxide	$\left\{\begin{array}{l}TlO\\O\\TlO\end{array}\right.$.
Thallous sulphide	STl_2.
Thallous nitrate	NO_2Tlo.
Thallous sulphate	SO_2Tlo_2.
Thallous carbonate	$COTlo_2$.
Thallic oxyhydrate	$TlOHo$.

CHAPTER XXI.

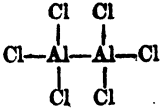
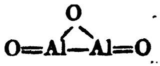
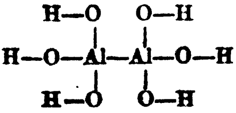
TETRAD ELEMENTS.

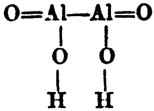
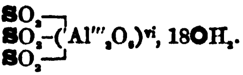
SECTION II.

ALUMINIUM, Al.

Atomic weight =27·5. *Molecular weight unknown. Specific gravity* 2·6. *Fuses at about* 450°. *Atomicity* $^{\text{iv}}$, *but is always a pseudo-triad. Evidence of Atomicity—analogy with iron and chromium.*

Annexed are the names and probable constitutional formulæ of the most important compounds of this metal :—

Aluminic chloride ...	$'Al'''_2Cl_6$ or $\left\{\begin{matrix} AlCl_3 \\ AlCl_3 \end{matrix}\right.$.	 $\begin{array}{ccccccc} & & Cl & & Cl & & \\ & & \vert & & \vert & & \\ Cl & — & Al & — & Al & — & Cl \\ & & \vert & & \vert & & \\ & & Cl & & Cl & & \end{array}$
Aluminic oxide	Al_2O_3 or $\left\{\begin{matrix} AlO \\ AlO \end{matrix}\right\}O$.	
Aluminic hydrate. (*Gibbsite.*)	Al_2Ho_6 or $\left\{\begin{matrix} AlHo_3 \\ AlHo_3 \end{matrix}\right.$.	 $\begin{array}{ccccccccc} & & H—O & & O—H & & & & \\ & & \vert & & \vert & & & & \\ H—O & — & Al & — & Al & — & O—H & & \\ & & \vert & & \vert & & & & \\ & & H—O & & O—H & & & & \end{array}$

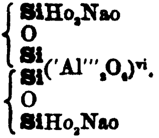
Aluminic oxydihydrate. (*Diaspore.*)	$\mathbf{Al}_2O_2Ho_2$ or $\left\{\begin{matrix}\mathbf{Al}OHo\\ \mathbf{Al}OHo\end{matrix}\right.$.
	
Aluminic sulphide ...	$\mathbf{Al}_2S''_3$ or $\left\{\begin{matrix}\mathbf{Al}S''\\ \mathbf{Al}S''\end{matrix}\right.S''$.
Dipotassic aluminate	$\mathbf{Al}_2O_2Ko_2$ or $\left\{\begin{matrix}\mathbf{Al}OKo\\ \mathbf{Al}OKo\end{matrix}\right.$.
Magnesic aluminate. (*Spinelle.*)	$\mathbf{Al}_2O_2Mgo''$ or $\left\{\begin{matrix}\mathbf{Al}O\\ \mathbf{Al}O\end{matrix}\right.Mgo''$.
Aluminic sulphate ..	$\mathbf{S}_3O_6('Al'''_2O_6)^{vi}, 18\mathbf{O}H_2$ or
	 $\begin{matrix}\mathbf{S}O_2\\ \mathbf{S}O_2\\ \mathbf{S}O_2\end{matrix}\Big]\!-\!('Al'''_2O_6)^{vi}, 18\mathbf{O}H_2.$
Aluminic sulphate tetrahydrate. (*Aluminite.*) ...	$\mathbf{S}O_2('Al'''_2O_2Ho_4)'', 7\mathbf{O}H_2.$
Allophane	$\mathbf{Si}Ho_2('Al'''_2Ho_4O_2)''$, (2 or 4)$\mathbf{O}H_2$.
Prehnite	$\mathbf{Si}_3Ho_2Cao''_2('Al'''_2O_6)^{vi}$.
Topaz	$\mathbf{Si}_3('Al'''_2O_5F)^{v}('Al'''_2O_4F_2)^{iv}('Al'''_2O_4F)'$
Zoisite..................	$\mathbf{Si}_4Cao''_3\left(\left\{\begin{matrix}'Al'''_2O_5\\ O\\ 'Al'''_2O_5\end{matrix}\right.\right)^{x}$.
Spodumene	$\mathbf{Si}_{15}O_{15}Lio_6('Al'''_2O_6)^{vi}_4$.
Petalite	$\mathbf{Si}_{30}O_{45}Nao_2Lio_4('Al'''_2O_6)^{vi}_4$.

Alunite, alumstone …………	$\begin{matrix}\mathrm{SOHoKo}\\ \mathrm{SOHo_2}\\ \mathrm{SOHo_2}\\ \mathrm{SOHoKo}\end{matrix}\begin{matrix}\mathrm{('Al'''_2Ho_3O_3)'''}\\ \mathrm{('Al'''_2Ho_4O_2)''}\\ \mathrm{('Al'''_2Ho_3O_3)'''}\end{matrix}.$
Collyrite ……………	$\mathrm{SiHo_2('Al'''_2Ho_5O)_2,\ 4OH_2}.$
Dipotassic aluminic tetrasulphate. (*Common alum.*) …………	$\left.\begin{matrix}\mathrm{SO_2Ko}\\ \mathrm{SO_2}\\ \mathrm{SO_2}\\ \mathrm{SO_2Ko}\end{matrix}\right]\mathrm{('Al'''_2O_6)^{vi},\ 24OH_2}.$
Wörthite…… ……	$\begin{matrix}\mathrm{Si}\\ \mathrm{SiO}\\ \mathrm{Si}\end{matrix}\begin{matrix}\mathrm{('Al'''_2HoO_5)^{v}}\\ \mathrm{('Al'''_2HoO_5)^{v}}\end{matrix}.$
Miloschine …………	$\mathrm{SiHo_2('Al'''_2Ho_4O_2)''}.$
Porcelain clay of Passau…………	$\begin{matrix}\mathrm{SiHo_3}\\ \mathrm{SiHo_3}\end{matrix}\mathrm{('Al'''_2Ho_2O_4)^{iv}}.$
Cimolite, kaolin of Ellenbogen ……	$\left.\begin{matrix}\mathrm{SiHo_3}\\ \mathrm{SiO}\\ \mathrm{SiO}\\ \mathrm{SiHo_3}\end{matrix}\right]\mathrm{('Al'''_2O_6)^{vi}}.$
Agalmatolite ………	$\left.\begin{matrix}\mathrm{SiOHo}\\ \mathrm{SiO}\\ \mathrm{SiO}\\ \mathrm{SiOHo}\end{matrix}\right]\mathrm{('Al'''_2O_6)^{vi}}.$
Buchholzite, xenolite ……………	$\begin{matrix}\mathrm{Si}\\ \mathrm{Si}\\ \mathrm{Si}\end{matrix}\begin{matrix}\mathrm{('Al'''_2O_6)^{vi}}\\ \mathrm{('Al'''_2O_6)^{vi}}\end{matrix}.$
Porcelain clay.. ……	$\left\{\begin{matrix}\mathrm{SiHo}\\ \mathrm{O}\\ \mathrm{SiHo}\end{matrix}\right]\mathrm{('Al'''_2Ho_2O_4)^{iv}}.$
Andalusite, chiastolite, cyanite, fibrolite, sillimanite ……………	$\mathrm{SiO('Al'''_2O_4)''}.$
Wernerite …………	$\mathrm{Si_2Cao''('Al'''_2O_6)^{vi}}.$

Saponite …………… $Si_7Mgo''_6Ho_{10}('Al'''_2O_6)^{vi}$.

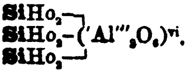
Lepidolite ………… $Si_8O_6Ko_2Lio_4('Al'''_2O_6)^{vi}_2('Al'''_2F_4O_2)''$.

Analcime …………



$$\left\{\begin{matrix} SiHo_2Nao \\ O \\ Si \end{matrix}\right. ('Al'''_2O_6)^{vi}. \quad \left\{\begin{matrix} Si \\ O \\ SiHo_2Nao \end{matrix}\right.$$

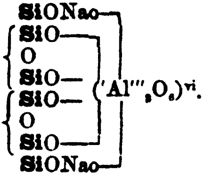
Razoumoffskin ……



$$\left.\begin{matrix} SiHo_2- \\ SiHo_2- \\ SiHo_2- \end{matrix}\right\} ('Al'''_2O_6)^{vi}.$$

Malthacite ………… $Si_8O_{11}Ho_4('Al'''_2O_6)^{vi}$.

Albite ………………



$$\left.\begin{matrix} SiONao- \\ \left\{\begin{matrix} SiO- \\ O \\ SiO- \end{matrix}\right. \\ \left\{\begin{matrix} SiO- \\ O \\ SiO- \end{matrix}\right. \\ SiONao- \end{matrix}\right\} ('Al'''_2O_6)^{vi}.$$

CHAPTER XXII.

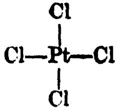
TETRAD ELEMENTS.

SECTION III.

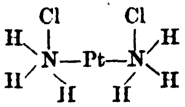
PLATINUM, Pt.

Atomic weight = 197·4. *Molecular weight unknown. Sp. gr.* 21·5. *Atomicity* '' *and* ^{iv}.

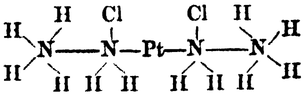
The following compounds will serve to illustate the atomicity and general character of this metal :—

Platinous chloride ...	$\mathbf{Pt}Cl_2$.	Cl—Pt—Cl
Platinic chlorido......	$\mathbf{Pt}Cl_4$.	 $$\begin{matrix} & Cl & \\ & \vert & \\ Cl— & Pt & —Cl \\ & \vert & \\ & Cl & \end{matrix}$$
Platinous oxide	$\mathbf{Pt}O$.	
Platinic oxide	$\mathbf{Pt}O_2$.	O=Pt=O
Platinous hydrate ...	$\mathbf{Pt}Ho_2$.	
Platinic hydrate......	$\mathbf{Pt}Ho_4$.	

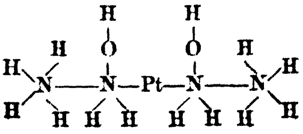
Platinosodiammonic dichloride. (*Green salt of Magnus.*) .... $\left\{\begin{matrix} \mathbf{N}H_3Cl \\ Pt'' \\ \mathbf{N}H_3Cl \end{matrix}\right.$.

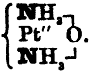
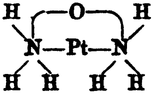


White compound of Reiset .... $\left\{\begin{matrix} \mathbf{N}H_2(N^{v}H_4)Cl \\ Pt'' \\ \mathbf{N}H_2(N^{v}H_4)Cl \end{matrix}\right.$.



Platoso-diammon diammonic dihydrate $\left\{\begin{matrix} \mathbf{N}H_2(N^{v}H_4)Ho \\ Pt'' \\ \mathbf{N}H_2(N^{v}H_4)Ho \end{matrix}\right.$.



Platosodiammonic oxide	 $\left\{\begin{matrix} NH_3 \\ Pt'' \\ NH_3 \end{matrix}\right\} O.$	
Platinous sulphide ...	PtS''.	
Platinic sulphide ...	PtS''_2.	

CHAPTER XXIII.

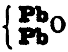
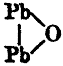
TETRAD ELEMENTS.

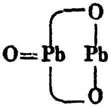
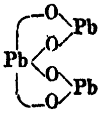
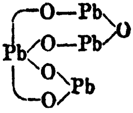
SECTION IV.

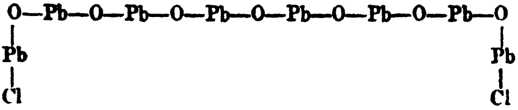
LEAD, Pb.

Atomic weight =207. *Molecular weight unknown. Sp. gr.* 11·445. *Fuses at* 335°. *Boils at a white heat. Atomicity* $''$ *and* ^{iv}. *Also sometimes a pseudo-triad.*

The following list contains the names and probable constitutional formulæ of the most interesting compounds of this metal :—

Plumbic chloride ...	$PbCl_2$.	Cl—Pb—Cl
Plumbous oxide	 $\left\{\begin{matrix} Pb \\ Pb \end{matrix}\right. O$	
Plumbic oxide. (*Litharge.*)	PbO.	Pb=O
Plumbic peroxide. (*Plattnerite.*) ...	PbO_2.	O=Pb=O

Red leads.	Diplumbic trioxide	$\mathbf{Pb}OPbo''$.	
	Triplumbic tetroxide ...	$\mathbf{Pb}Pbo''_2$.	
	Tetraplumbic pentoxide	$\mathbf{Pb}Pbo''(Pb''_2O_3)''$.	
	Plumbic chlorohydrate	$\mathbf{Pb}HoCl$.	Cl—Pb—O—H
	Plumbic hydrate......	$\mathbf{Pb}Ho_2$.	H—O—Pb—O—H
	Diplumbic oxydihydrate	$\left\{\begin{array}{l}\mathbf{Pb}Ho\\ O\\ \mathbf{Pb}Ho\end{array}\right.$.	H—O—Pb—O—Pb—O—H
	Diplumbic oxydichloride. (*Matlockite.*)	$\left\{\begin{array}{l}\mathbf{Pb}Cl\\ O\\ \mathbf{Pb}Cl\end{array}\right.$.	Cl—Pb—O—Pb—Cl
	Diplumbic oxychlorohydrate	$\left\{\begin{array}{l}\mathbf{Pb}Ho\\ O\\ \mathbf{Pb}Cl\end{array}\right.$.	Cl—Pb—O—Pb—O—H
	Octoplumbic heptoxydichloride ...	$\mathbf{Pb}_8O_7Cl_2$.	



$$\begin{array}{ccc} \mathrm{O} & \mathrm{{-}Pb{-}O{-}Pb{-}O{-}Pb{-}O{-}Pb{-}O{-}Pb{-}O{-}Pb{-}} & \mathrm{O} \\ | & & | \\ \mathrm{Pb} & & \mathrm{Pb} \\ | & & | \\ \mathrm{Cl} & & \mathrm{Cl} \end{array}$$

Triplumbic oxydichloride. (*Mendipite.*) .... $\left\{\begin{array}{l}\mathbf{Pb}Cl\\ O\\ Pb''\\ O\\ \mathbf{Pb}Cl\end{array}\right.$. Cl—Pb—O—Pb—O—Pb—Cl

Plumbic sulphide. (*Galena.*) ...... $\mathbf{Pb}S''$.

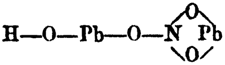
Diplumbic sulphodichloride .... $\left\{\begin{array}{l}\mathbf{Pb}Cl\\ S''\\ \mathbf{Pb}Cl\end{array}\right.$. Cl—Pb—S—Pb—Cl

Plumbic sulphate. (*Lead vitriol.*)... $\mathbf{S}O_2Pbo''$.

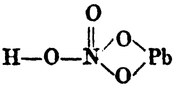
Plumbic nitrite ...... $\left\{\begin{array}{l}\mathbf{N}O\\ Pbo''.\\ \mathbf{N}O\end{array}\right.$

Plumbic nitrite hydrate .... $\mathbf{N}O(Pb''HoO)$.

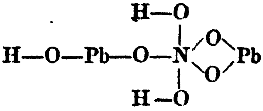
Diplumbic nitrite-hydrate .... $\mathbf{N}Pbo''(Pb''HoO)$.



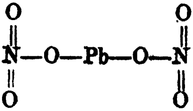
Hydric plumbic nitrate............... $\mathbf{N}OHoPbo''$.



Dihydric diplumbic nitrate hydrate .... $\mathbf{N}Ho_2Pbo''(PbHoO)$.

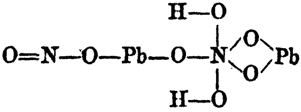


Plumbic nitrate ...... $\left\{\begin{matrix}NO_2\\ Pbo''.\\ NO_2\end{matrix}\right.$

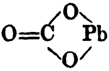


$$\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{N}}-O-Pb-O-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{N}}$$

Dihydric diplumbic nitrate nitrite. (*Basic hyponitrate of lead.*) ... $\left.\right\}\begin{matrix}NHo_2Pbo''\\ Pbo''\\ NO\end{matrix}$.

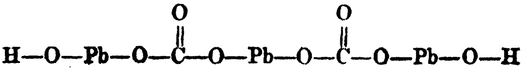


Plumbic carbonate. (*Lead spar, white lead ore.*) .... $\left.\right\}$ $COPbo''$.



Yorke's diplumbic dihydrate carbonate .... $\left.\right\}$ $CO(OPb''Ho)_2$.

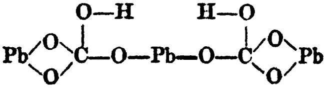
Triplumbic dihydrate dicarbonate. (*White lead.*) ... $\left.\right\}$ $\left.\begin{matrix}CO(OPbHo)\\ CO(OPbHo)\end{matrix}\right\}Pbo''$;



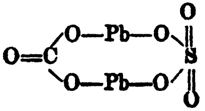
$$H-O-Pb-O-\overset{\overset{O}{\|}}{C}-O-Pb-O-\overset{\overset{O}{\|}}{C}-O-Pb-O-H$$

or

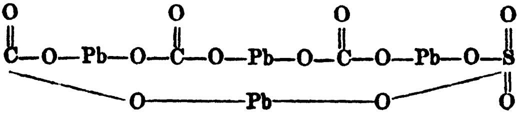
Dihydric triplumbic dicarbonate $\left.\right\}$ $\left.\begin{matrix}CHoPbo''\\ CHoPbo''\end{matrix}\right\}Pbo''$.



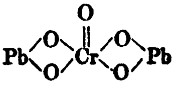
Diplumbic sulphate carbonate. (*Lanarkite.*) ... $\left.\right\}$ $\begin{matrix}\mathbf{CO}\\\mathbf{SO_2}\end{matrix}Pbo''_2.$



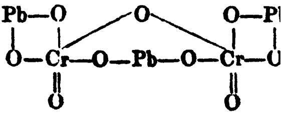
Tetraplumbic tricarbonate sulphate. (*Leadhillite.*) .... $\left.\right\}$ $\left.\begin{matrix}\mathbf{CO}\\\mathbf{CO}\\\mathbf{CO}\\\mathbf{SO_2}\end{matrix}\begin{matrix}Pbo''\\Pbo''\\Pbo''\end{matrix}\right]Pbo''.$



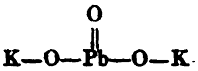
Diplumbic chromate. $\mathbf{Cr}OPbo''_2.$



Triplumbic dichromate ...... $\left\{\begin{matrix}\mathbf{Cr}OPbo''\\Pbo''\\\mathbf{Cr}OPbo''\end{matrix}\right]O.$



Dipotassic plumbate ...... $\left.\right\}$ $\mathbf{Pb}OKo_2.$



CHAPTER XXIV.

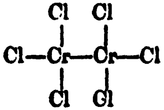
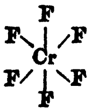
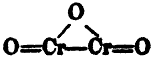
HEXAD ELEMENTS.

SECTION IV.

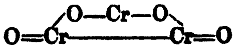
CHROMIUM, Cr.

Atomic weight =52·5. *Molecular weight unknown. Sp. gr.* 7·0146. *Atomicity* $''$, ^{iv}, and ^{vi}; *also a pseudo-triad* (*and a pseud-octad*). *Evidence of atomicity*:—*See the annexed compounds.*

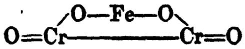
The following are the names and probable constitutional formulæ of the most important compounds of this metal:—

Chromous chloride	$CrCl_2$.	Cl—Cr—Cl
Chromic chloride …	$\left\{\begin{matrix}CrCl_3\\CrCl_3\end{matrix}\right.$.	 $\begin{array}{ccccc} & Cl & & Cl & \\ & \mid & & \mid & \\ Cl- & Cr & - & Cr & -Cl \\ & \mid & & \mid & \\ & Cl & & Cl & \end{array}$
Chromic perfluoride	CrF_6.	
Chromous oxide ……	CrO.	
Chromous hydrate …	$CrHo_2$.	
Chromic oxide ……	$\left\{\begin{matrix}CrO\\CrO\end{matrix}\right.O$.	

Chromous dichromic tetroxide .... $\left\{\begin{matrix}\mathbf{Cr}O\\ \mathbf{Cr}O\end{matrix}\right.Cro''.$

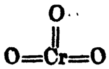


Dichromic ferrous tetroxide. (*Chrome iron ore.*) $\left\{\begin{matrix}\mathbf{Cr}O\\ \mathbf{Cr}O\end{matrix}\right.Feo''.$



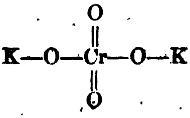
Chromic dioxide...... $\mathbf{Cr}O_2$.

Chromic anhydride. . $\mathbf{Cr}O_3$.



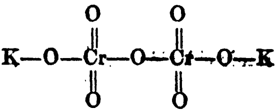
$$O{=}\overset{\overset{O}{\|}}{Cr}{=}O$$

Dipotassic chromate. (*Normal potassic chromate.*) .... $\mathbf{Cr}O_2Ko_2$.



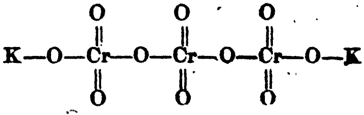
$$K{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}{-}O{-}K$$

Dipotassic dichromate. (*Potassic bichromate.*)...... $\left\{\begin{matrix}\mathbf{Cr}O_2Ko\\ O\\ \mathbf{Cr}O_2Ko\end{matrix}\right.$.

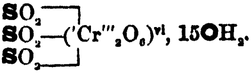
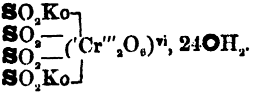
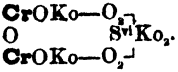


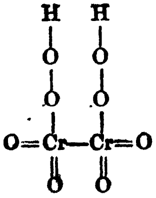
$$K{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}{-}O{-}K$$

Dipotassic trichromate. (*Potassic terchromate.*) $\left\{\begin{matrix}\mathbf{Cr}O_2Ko\\ O\\ \mathbf{Cr}O_2\\ O\\ \mathbf{Cr}O_2Ko\end{matrix}\right.$

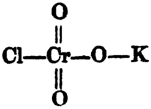


$$K{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}{-}O{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}{-}O{-}K$$

Dipotassic chromous disulphate	$\left.\begin{matrix}\mathbf{S}O_2Ko\\\mathbf{S}O_2Ko\end{matrix}\right.Cro'', 6\mathbf{O}H_2.$
Chromic sulphate ...	 $\left\{\begin{matrix}\mathbf{S}O_2\\\mathbf{S}O_2\\\mathbf{S}O_2\end{matrix}\right]\!-({}'Cr'''_2O_6)^{vi}, 15\mathbf{O}H_2.$
Dipotassic chromic sulphate (*Potassium chrome alum.*)	 $\left.\begin{matrix}\mathbf{S}O_2Ko\\\mathbf{S}O_2\\\mathbf{S}O_2\\\mathbf{S}O_2Ko\end{matrix}\right]\!-({}'Cr'''_2O_6)^{vi}, 24\mathbf{O}H_2.$
Chromic nitrate	$\mathbf{N}_6O_{12}({}'Cr'''_2O_6)^{vi}.$
Octochromic carbonate dihydrate ...	$\mathbf{C}({}'Cr'''_8O_{13}Ho_2)^{iv}.$
Plumbic chromate. (*Red lead ore, crocoisite.*)	$\mathbf{Cr}O_2Pbo''.$
Sulphochromic acid. (*Dihydric sulphate chromate.*)	$\left\{\begin{matrix}\mathbf{S}O_2Ho\\O\\\mathbf{Cr}O_2Ho\end{matrix}\right..$
Tetrapotassic dichromosulphate	 $\left\{\begin{matrix}\mathbf{Cr}OKo—O_2\\O\\\mathbf{Cr}OKo—O_2\end{matrix}\right]S^{vi}Ko_2.$
Chromic hydrate	$\left\{\begin{matrix}\mathbf{Cr}Ho_3\\\mathbf{Cr}Ho_3\end{matrix}\right., 6\mathbf{O}H_2.$
Perchromic acid	$\left\{\begin{matrix}\mathbf{Cr}O_2(OHo)\\\mathbf{Cr}O_2(OHo)\end{matrix}\right..$

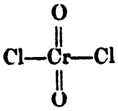


Potassic chlorochromate................ CrO_2ClKo.



$Cl-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}-O-K$

Chromic dichlorodioxide. (*Chlorochromic acid.*) ...... CrO_2Cl_2.



$Cl-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Cr}}-Cl$

Chromous sulphide... CrS''.

Dichromic trisulphide .... $\left\{\begin{matrix}CrS''\\CrS''\end{matrix}\right.S''$.

Sulphochromic anhydride .... CrS''_3.

MANGANESE, Mn.

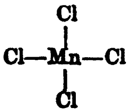
Atomic weight =55. *Molecular weight unknown. Sp. gr.* 7 *to* 8. *Atomicity* $''$, ^{iv}, *and* ^{vi}; *also a pseudo-triad and a pseud-octad. Evidence of atomicity*:—

Manganic perfluoride .... $Mn^{vi}F_6$.
Analogy with chromium.

The following are some of the more important compounds of this element:—

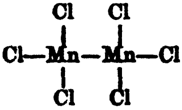
Manganous chloride. $MnCl_2$. $Cl-Mn-Cl$

Manganic chloride... $MnCl_4$.



$Cl-\underset{\underset{Cl}{|}}{\overset{\overset{Cl}{|}}{Mn}}-Cl$

Dimanganic hexachloride .... $\left\{\begin{matrix}MnCl_3\\MnCl_3\end{matrix}\right.$.

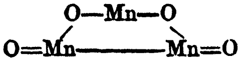


$Cl-\underset{\underset{Cl}{|}}{\overset{\overset{Cl}{|}}{Mn}}-\underset{\underset{Cl}{|}}{\overset{\overset{Cl}{|}}{Mn}}-Cl$

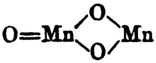
Manganous hydrate.. $\mathbf{Mn}Ho_2$. $H—O—Mn—O—H$

Manganous oxide ... $\mathbf{Mn}O$. $Mn{=}O$

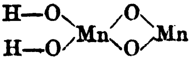
Manganous dimanganic tetroxide. (*Hausmannite.*) ... $\begin{matrix}\mathbf{Mn}O\\ \mathbf{Mn}O\end{matrix}Mno''$.



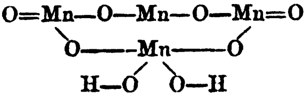
Dimanganic trioxide. (*Braunite.*) .... $\mathbf{Mn}OMno''$.



Dimanganic dioxydihydrate. (*Manganite.*) .... $\mathbf{Mn}Ho_2Mno''$.

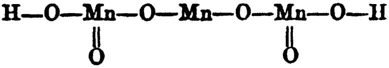


Varvicite.............. $\left\{\begin{matrix}\mathbf{Mn}O\\ Mno''\ (Mn^{iv}O_2Ho_2)''\\ \mathbf{Mn}O\end{matrix}\right.$



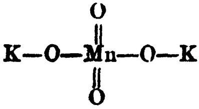
Manganic oxide. (*Pyrolusite.*) ...... $\mathbf{Mn}O_2$.

Manganous dimanganic tetroxide dihydrate. (*Psilomelane, Hartmangan*). $\begin{matrix}\mathbf{Mn}OHo\\ \mathbf{Mn}OHo\end{matrix}Mno''$.

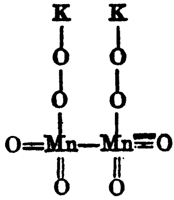


$$H—O—\underset{\substack{\| \\ O}}{Mn}—O—Mn—O—\underset{\substack{\| \\ O}}{Mn}—O—H$$

Dipotassic manganate $\mathbf{Mn}O_2Ko_2$.



$$K—O—\overset{\substack{O \\ \|}}{\underset{\substack{\| \\ O}}{Mn}}—O—K$$

Dipotassic permanganate	$\left\{\begin{matrix}MnO_2(OKo)\\MnO_2(OKo)\end{matrix}\right.$. 
Manganous sulphide. (*Manganese blende.*)...	MnS.
Disulphopotassic trimanganous disulphide	$\begin{matrix}MnKs\\MnKs\end{matrix}Mns''$
Manganous carbonate. (*Manganese spar.*)	$COMno''$.
Dihydric manganous sulphate ...	$SOHo_2Mno''$, (3, 4, or $6OH_2$).
Dipotassic manganous disulphate.	$\begin{matrix}SO_2Ko\\SO_2Ko\end{matrix}Mno''$, $6OH_2$.
Aluminic manganous tetrasulphate. (*Manganese aluminium alum.*)	$\begin{matrix}SO_2\\ \left\{\begin{matrix}SO_2\\Mno''(Al'''_2O_6)^{vi}\\SO_2\end{matrix}\right.\\SO_2\end{matrix}$, $24OH_2$.
Dimanganic trisulphate	$\left.\right\}\begin{matrix}SO_2\\SO_2\\SO_2\end{matrix}(Mn'''_2O_6)^{vi}$.
Dipotassic dimanganic tetrasulphate. (*Potassium manganese alum.*)	$\begin{matrix}SO_2Ko\\SO_2\\SO_2\\SO_2Ko\end{matrix}(Mn'''_2O_6)^{vi}$, $24OH_2$.

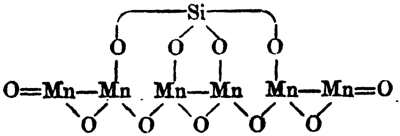
Manganous hydric phosphate ..... $\mathbf{P}OHoMno''$.

Manganous silicate. (*Siliciferous manganese, red manganese, rother Mangankiesel, Rothbraunsteinerz.*) .... $\mathbf{Si}OMno''$.

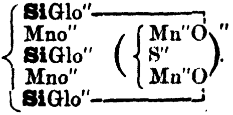
Dimanganous silicate. (*Tephroite.*) .... $\mathbf{Si}Mno''_2$.

Dihydric dimanganous silicate dihydrate. (*Schwarzer Mangankiesel.*) .... $\mathbf{Si}Ho_2(OMnHo)_2$.

Hexamanganic monosilicate. (*Heterocline.*) .... $\mathbf{Si}('Mn'''_6O_{11})^{iv}$.

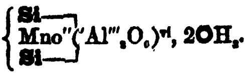


Triglucinic tetramanganous trisilicate sulphide. (*Helvine.*)



$$\left\{\begin{array}{l}\mathbf{Si}Glo''\\ Mno''\\ \mathbf{Si}Glo''\\ Mno''\\ \mathbf{Si}Glo''\end{array}\right. \left(\left\{\begin{array}{l}Mn''O\\ S''\\ Mn''O\end{array}\right.\right)''$$

Aluminic manganous disilicate.



$$\left\{\begin{array}{l}\mathbf{Si}\\ Mno''\\ \mathbf{Si}\end{array}\right.('Al'''_2O_6)^{vi},\ 2\mathbf{O}H_2.$$

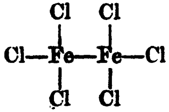
IRON, Fe.

Atomic weight =56. *Molecular weight unknown. Sp. gr.* 7·8. *Atomicity* $''$, ^{iv}, *and* ^{vi}. *Evidence of atomicity—analogy with chromium.*

The following is a list of the chief compounds of iron:—

Ferrous chloride ... $\mathbf{Fe}Cl_2$. Cl—Fe—Cl

Ferric chloride ...... $\left\{\begin{matrix}\mathrm{FeCl_3.}\\\mathrm{FeCl_3.}\end{matrix}\right.$

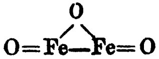


$$\mathrm{Cl—\underset{\substack{|\\Cl}}{\overset{\substack{Cl\\|}}{Fe}}—\underset{\substack{|\\Cl}}{\overset{\substack{Cl\\|}}{Fe}}—Cl}$$

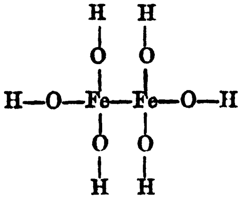
Ferrous oxide......... $\mathrm{FeO.}$ $\mathrm{Fe{=}O}$

Ferrous hydrate ...... $\mathrm{FeHo_2.}$ $\mathrm{H—O—Fe—O—H}$

Ferric oxide. (*Red hæmatite, micaceous iron, oligist, specular iron,* or *iron glance.*) ...... $\left\{\begin{matrix}\mathrm{FeO}\\\mathrm{FeO}\end{matrix}\right.\mathrm{O.}$

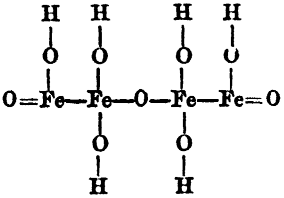


Ferric hydrate ...... $\left\{\begin{matrix}\mathrm{FeHo_3}\\\mathrm{FeHo_3}\end{matrix}\right.$



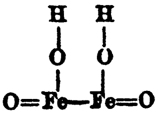
$$\mathrm{H—O—\underset{\substack{|\\O\\|\\H}}{\overset{\substack{H\\|\\O\\|}}{Fe}}—\underset{\substack{|\\O\\|\\H}}{\overset{\substack{H\\|\\O\\|}}{Fe}}—O—H}$$

Tetraferric trioxy-hexahydrate. (*Compact brown iron ore, fibrous brown iron ore* or *brown hæmatite.*) .... $\left\{\begin{matrix}\mathrm{FeOHo}\\\mathrm{FeHo_2}\\\mathrm{O}\\\mathrm{FeHo_2}\\\mathrm{FeOHo}\end{matrix}\right.$.



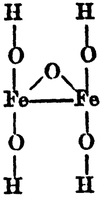
$$\mathrm{O{=}\overset{\substack{H\\|\\O\\|}}{Fe}—\underset{\substack{|\\O\\|\\H}}{\overset{\substack{H\\|\\O\\|}}{Fe}}—O—\underset{\substack{|\\O\\|\\H}}{\overset{\substack{H\\|\\O\\|}}{Fe}}—\overset{\substack{H\\|\\O\\|}}{Fe}{=}O}$$

Ferric dioxy-dihydrate. (*Needle iron ore, brown iron ore.*) $\left\{\begin{matrix}\mathrm{FeOHo}\\\mathrm{FeOHo}\end{matrix}\right.$

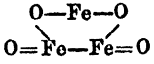


$$\mathrm{O{=}\overset{\substack{H\\|\\O\\|}}{Fe}—\overset{\substack{H\\|\\O\\|}}{Fe}{=}O}$$

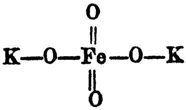
Ferric oxy-tetrahydrate. (Another variety of *brown iron ore.*) $\left\{\begin{matrix}\mathbf{Fe}Ho_2\\ \mathbf{Fe}Ho_2\end{matrix}\right.O.$



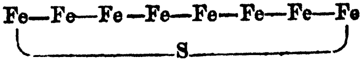
Ferrous ferric tetroxide. (*Magnetic iron ore.*) $\left\{\begin{matrix}\mathbf{Fe}O\\ \mathbf{Fe}O\end{matrix}\right.Feo''.$



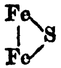
Potassic ferrate ...... $\mathbf{Fe}O_2Ko_2.$



Octoferrous sulphide ${}^{xiv}(\mathbf{Fe}_8)''S''.$



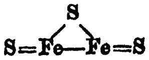
Diferrous sulphide... $\left\{\begin{matrix}\mathbf{Fe}\\ \mathbf{Fe}\end{matrix}\right.S''.$



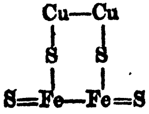
Ferrous sulphide ... $\mathbf{Fe}S''.$

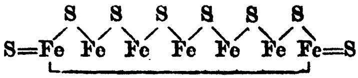
Ferric disulphide. (*Iron pyrites, martial pyrites.*) $\mathbf{Fe}S''_2.$ S=Fe=S

Diferric trisulphide $\left\{\begin{matrix}\mathbf{Fe}S''\\ \mathbf{Fe}S''\end{matrix}\right.S''.$



Diferric cuprous tetrasulphide. (*Copper pyrites.*) $\left\{\begin{matrix}\mathbf{Fe}S\\ \mathbf{Fe}S\end{matrix}\right.(Cu_2S''_2)''.$

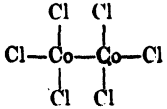
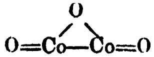
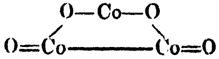
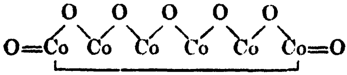
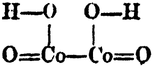


Ferric nitrate	$\mathbf{N}_6O_{12}({}'\mathrm{Fe}'''_2O_6)^{vi}$.
Hexahydric diferric diphosphate dihydrate.	$\mathbf{P}_2Ho_6({}'\mathrm{Fe}''''_2Ho_2O_4)^{iv}$.
Ferrous sulphate ...	$\mathbf{S}OHo_2Feo'', 6\mathbf{O}H_2$.
Dipotassic ferrous disulphate	$\left\{\begin{matrix}\mathbf{S}O_2Ko\\ Feo''\\ \mathbf{S}O_2Ko\end{matrix}\right., 6\mathbf{O}H_2$.
Ferric sulphate. (*Coquimbite.*) ...	$\left.\begin{matrix}\mathbf{S}O_2\\ \mathbf{S}O_2\\ \mathbf{S}O_2\end{matrix}\right]-({}'\mathrm{Fe}'''_2O_6)^{vi}, 9\mathbf{O}H_2$.
Dipotassic ferric tetrasulphate. (*Potassium iron alum.*)	$\left.\begin{matrix}\mathbf{S}O_2Ko\\ \mathbf{S}O_2\\ \mathbf{S}O_2\\ \mathbf{S}O_2Ko\end{matrix}\right]({}'\mathrm{Fe}'''_2O_6)^{vi}, 24\mathbf{O}H_2$.
Tetraferric sulphate..	$\mathbf{S}O_2({}'\mathrm{Fe}'''_4O_7)'', 6\mathbf{O}H_2$.
Tetrahydric tetraferric sulphate octahydrate. (*Vitriol ochre.*)	$\mathbf{S}Ho_4({}'\mathrm{Fe}'''_4OHo_6O_2)''$.
Heptaferric octosulphide. (*Magnetic pyrites.*) ...	$\mathbf{Fe}_7S''_8$. 
Ferrous carbonate. (*Spathic iron ore.*)	$\mathbf{C}OFeo''$.
Ferrous nitrate ...	$\left\{\begin{matrix}\mathbf{N}O_2\\ Feo''\\ \mathbf{N}O_2\end{matrix}\right.$.

COBALT, Co.

Atomic weight* = 58·8. *Molecular weight unknown.* *Sp. gr.* 8·5. *Atomicity* '', ^{iv}, *and* ^{vi}? *also a pseudo-triad.

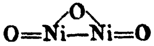
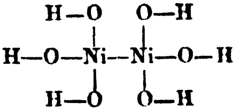
For evidence of the atomicity of cobalt see the following list of the chief compounds of this metal:—

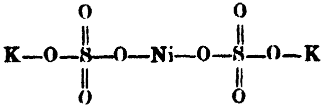
Cobaltous chloride ...	$\mathbf{Co}Cl_2$.	Cl—Co—Cl
Cobaltic chloride ...	$\left\{\begin{matrix}\mathbf{Co}Cl_3\\ \mathbf{Co}Cl_3\end{matrix}\right.$.	
Cobaltous oxide	$\mathbf{Co}O$.	Co=O
Cobaltic oxide........	$\left\{\begin{matrix}\mathbf{Co}O\\ \mathbf{Co}O\end{matrix}\right.O$.	
Cobaltous dicobaltic tetroxide	$\left\{\begin{matrix}\mathbf{Co}O\\ \mathbf{Co}O\end{matrix}\right.Coo''$.	
Hexacobaltic heptoxide	$\mathbf{Co}_6O_7$.	
Cobaltous hydrate	$\mathbf{Co}Ho_2$	H—O—Co—O—H
Cobaltic oxy-dihydrate	$\left\{\begin{matrix}\mathbf{Co}OHo\\ \mathbf{Co}OHo\end{matrix}\right.$.	
Cobaltous sulphide...........	$\mathbf{Co}S''$.	
Cobaltic sulphide. (*Cobalt pyrites.*)	$\left\{\begin{matrix}\mathbf{Co}S''\\ \mathbf{Co}S''\end{matrix}\right.S''$.	

NICKEL, Ni.

Atomic weight =58·8. *Molecular weight unknown.* *Sp. gr.* 8·7.
Atomicity $''$, $^{\text{iv}}$, *and* $^{\text{vi}}$? *also a pseudo-triad.*

Annexed is a list of the names and probable constitutional formulæ of the chief compounds of this metal:—

Nickelous chloride.........	$\mathbf{Ni}Cl_2$.	Cl—Ni—Cl
Nickelous oxide	$\mathbf{Ni}O$.	
Nickelic oxide....	$\left\{\begin{array}{l}\mathbf{Ni}O\\ \mathbf{Ni}O\end{array}\right.O$.	
Nickelic hydrate...........	$\left\{\begin{array}{l}\mathbf{Ni}Ho_3\\ \mathbf{Ni}Ho_3\end{array}\right.$.	 $\begin{array}{ccc} & H—O & O—H \\ & \vert & \vert \\ H—O— & Ni— & Ni—O—H \\ & \vert & \vert \\ & H—O & O—H \end{array}$
Nickelous hydrate	$\mathbf{Ni}Ho_2$.	
Dinickelous sulphide	$\left\{\begin{array}{l}\mathbf{Ni}\\ \mathbf{Ni}\end{array}\right.S''$.	
Nickelous sulphide. (*Capillary pyrites, hair nickel.*).........	$\mathbf{Ni}S''$.	
Nickelic disulphide......	$\mathbf{Ni}S''_2$.	S=Ni=S
Dihydric nickelous sulphate	$\mathbf{S}OHo_2Nio''$, $6\mathbf{O}H_2$.	
Dipotassic nickelous disulphate	$\left\{\begin{array}{l}\mathbf{S}O_2Ko\\ Nio''\\ \mathbf{S}O_2Ko\end{array}\right.$, $6\mathbf{O}H_2$.	



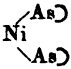
$$\begin{array}{ccccc} & O & & O & \\ & \Vert & & \Vert & \\ K—O— & S & —O—Ni—O— & S & —O—K \\ & \Vert & & \Vert & \\ & O & & O & \end{array}$$

Nickelous nitrate …… .. $\left\{\begin{array}{l}\mathbf{N}O_2\\ Nio'',\ 6\mathbf{O}H_2.\\ \mathbf{N}O_2\end{array}\right.$

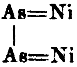
Dihydric pentanickelous dicarbonate tetrahydrate ………………… $\left\{\begin{array}{l}\mathbf{C}Ho(Ni''OHo)_2\\ Nio''\qquad\qquad\quad ,\ 2\mathbf{O}H_2.\\ \mathbf{C}Ho(Ni''OHo)_2\end{array}\right.$

Dipotassic nickelous tricarbonate …………… $\left\{\begin{array}{l}\mathbf{C}OKo\\ O\\ \mathbf{C}Nio'',\ 10\mathbf{O}H_2\\ O\\ \mathbf{C}OKo\end{array}\right.$

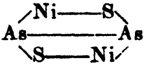
Nickelous diarsenide. (*Arsenical nickel.*) $\left.\right\}$ $\begin{array}{l}{}''\mathbf{As}'\\ {}''\mathbf{As}'\end{array}Ni''.$



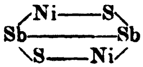
Dinickelous diarsenide. (*Kupfernickel.*) …… $\left\{\begin{array}{l}\mathbf{As}Ni\\ \mathbf{As}Ni\end{array}\right.$


$\begin{array}{l}As{=}Ni\\ \vert\\ As{=}Ni\end{array}$

Nickelous sulpharsenide. (*Grey nickel ore, nickel glance.*) ……………… $\left\{\begin{array}{l}\mathbf{As}\\ \mathbf{As}\end{array}\right.(NiS)''_2$



Nickelous sulphantimonide. (*Antimony nickel glance.*)……… $\left\{\begin{array}{l}\mathbf{Sb}\\ \mathbf{Sb}\end{array}\right.(NiS)''_2.$



INDEX TO VOL. I.

Absolute atomicity, 21.
Acid, arsenic, 128.
,, arsenious, 127.
,, boracic, 56.
,, boric, 56.
,, boric, monobasic, 55.
,, boric, tribasic, 55.
,, bromic, 91.
,, carbonic, 60.
,, chlorhydric, 39.
,, chloric, 50.
,, chlorochromic, 192.
,, chlorous, 49.
,, disulphodithionic, 77, 85.
,, disulphuretted hyposulphuric, 85.
,, dithionic, 76, 84.
,, graphitic, 58.
,, hydriodic, 93.
,, hydrobromic, 89.
,, hydrochloric, 39.
,, hydrofluoboric, 55.
,, hydrofluoric, 99.
,, hydroselenic, 86.
,, hydrosulphuric, 72.
,, hypobromous, 91.
,, hypochlorous, 48.
,, hypophosphorous, 118.
,, hyposulphuric, 84.
,, hyposulphurous, 76, 77, 83.
,, iodic, 95, 96.
,, metabismuthic, 148.
,, metabismuthous, 146.
,, metaboric, 55, 56.
,, metantimonic, 139.
,, metantimonic, of Frémy, 140.
,, metantimonious, 137.
,, metaphosphoric, 119, 121.
,, metarsenic, 128.
,, metastannic, 109.
,, muriatic, 39.
,, nitric, 63.
,, nitric, manufacture of, 63.
,, nitrous, 63, 65.
,, orthantimonic, 139.
,, orthobismuthous, 147.
,, orthoboric, 56.
,, parantimonic, 139.
,, pentathionic 77, 85.
Acid, perchloric, 51.
,, perchromic, 191.
,, periodic, 95, 98.
,, phosphoric, 119, 122.
,, phosphoric, dodecabasic, 119.
,, phosphoric, hexabasic, 119.
,, phosphorous, 118, 120.
,, pyrantimonic, 140.
,, pyrarsenic, 128.
,, pyrophosphoric, 119, 122.
,, selenic, 87.
,, selenious, 87.
,, silicic, 104.
,, stannic, 108.
,, sulpharsenic, 129.
,, sulpharsenious, 129.
,, sulphhydric, 72.
,, sulphocarbonic, 75.
,, sulphochromic, 191.
,, sulphodithionic, 77, 84.
,, sulphuric, 76, 80.
,, sulphuric, manufacture of, 81.
,, sulphuric, Nordhausen, 76, 79.
,, sulphurous, 76.
,, tetrathionic, 77, 85.
,, thiosulphuric, 76, 83.
,, titanic, 110.
,, trisulphodithionic, 77, 85.
,, trisulphuretted hyposulphuric, 85.
,, trithionic, 77, 84.
Acid salts, definition of, 13.
Acids, anhydrous, definition of, 8, 10.
,, definition of, 8.
,, dibasic, definition of, 9.
,, formation from anhydrides, 44.
,, monobasic, definition of, 9.
,, nomenclature of, 9.
,, polybasic, definition of, 9.
Acids of antimony, 134.
,, of chlorine, 48.
,, of nitrogen, 63.
Action, chemical, modes of, 1
Active atomicity, 21.
Affinity, chemical, 4.
Agalmatolite, 181
Alabaster, 164.

Albite, 182.
Allophane, 180.
Allotropic oxygen, 42.
  „ phosphorus, 112.
  „ varieties of sulphur, 72.
Alum, common, 181.
  „ manganese aluminium, 194.
  „ potassium chrome, 191.
  „ potassium manganese, 194.
Aluminite, 180.
Aluminium, 179.
  Agalmatolite, 181.
  Albite, 182.
  Allophane, 180.
  Alum, 181.
  Alum, manganese aluminium, 194.
    „ potassium chrome, 191.
    „ potassium manganese, 194.
  Aluminate, dipotassic, 180.
    „ magnesic, 180.
  Aluminic calcic disilicate, 107.
    „ calcic trisilicate, 107.
    „ chloride, 179.
    „ hydrate, 179.
    „ manganous disilicate, 195.
    „ manganous tetrasulphate, 194.
    „ oxide, 179.
    „ oxydihydrate, 180.
    „ sulphate, 180.
    „ sulphate tetrahydrate, 180.
    „ sulphide, 180.
    „ tricalcic trisilicate, 107.
  Aluminite, 180.
  Alum-stone, 181.
  Alunite, 181.
  Analcime, 182.
  Andalusite, 181.
  Anorthite, 107.
  Atomicity of aluminium, 179.
  Buchholzite, 181.
  Chiastolite, 181.
  Cimolite, 181.
  Collyrite, 181.
  Compounds of aluminium, 179.
  Cyanite, 181.
  Diaspore, 180.
Aluminium (*continued*).
  Dihydric aluminic tetrasilicate, 107.
  Dipotassic aluminate, 180.
    „ aluminic hexasilicate, 107.
    „ aluminic tetrasulphate, 181.
  Emerald, 107.
  Felspar, 107.
  Fibrolite, 181.
  Gibbsite, 179.
  Grossularia, 107.
  Kaolin of Ellenbogen, 181.
  Labradorite, 107.
  Lepidolite, 182.
  Magnesic aluminate, 180.
  Malthacite, 182.
  Miloschine, 181.
  Orthose, 107.
  Porcelain clay, 181.
    „ clay of Passau, 181.
  Prehnite, 180.
  Pyrophyllite, 107.
  Razoumoffskin, 182.
  Saponite, 182.
  Sillimanite, 181.
  Spinelle, 180.
  Spodumene, 180.
  Topaz, 180.
  Triglucinic aluminic hexasilicate, 107.
  Wernerite, 181.
  Wörthite, 181.
  Xenolite, 181.
  Zoisite, 180.
Alum-stone, 181.
Alunite, 181.
Amidogen, 28.
Ammonia, 61, 68.
Ammonic chloride, 61, 68.
  „ salts, 69, 70.
Ammoniocupric carbonate, 176.
  „ sulphate, 175.
Ammonium, 69.
Ammonium amalgam, 70.
Ammonoxyl, 28.
Amorphous boron, 52.
  „ carbon, 58.
  „ phosphorus, 112.
Analcime, 182.
Anatase, 110.
Andalusite, 181.

Anhydride, antimonic, 136, 138.
,, antimonious, 136.
,, arsenic, 128.
,, arsenious, 127.
,, auric, 177.
,, bismuthic, 147.
,, boric, 55.
,, carbonic, 59.
,, chlorous, 47.
,, chromic, 190.
,, hypobromous, 91.
,, hypochlorous, 47.
,, iodic, 95.
,, nitric, 63, 64.
,, nitrous, 62, 65.
,, periodic, 95, 97.
,, phosphoric, 118, 121.
,, phosphorous, 118, 120.
,, selenious, 87.
,, silicic, 104, 106.
,, silicoformic, 102.
,, stannic, 108.
,, sulphantimonic, 143.
,, sulphantimonious, 141.
,, sulpharsenic, 129, 130.
,, sulpharsenious, 129.
,, sulphochromic, 192.
,, sulphuric, 76, 79.
,, sulphurous, 76, 77.
,, titanic, 110.
Anhydrides, conversion into acids, 44.
,, definition of, 8, 10.
Anhydrous acid, definition of, 8, 10.
Anorthite, 107.
Antimonic chloride, 130, 134.
,, oxytrichloride, 135.
,, sulphide, 143.
,, sulphotrichloride, 135.
,, tetretho-chloride, 130.
Antimonious amylide, 133.
,, anhydride, 136.
,, argentide, 132.
,, bromide, 133, 135.
,, chloride, 130, 133.
,, ethide, 133.
,, fluoride, 135.
,, hydride, 130.
,, iodide, 135.
,, oxide, 136.
,, oxybromide, 135.
,, oxychloride, 134.
,, oxydisulphide, 140.
,, oxyiodide, 135.
Antimonious sulphide, 141.
,, zincide, 133.
Antimoniuretted hydrogen, 132.
Antimony, 130.
,, amorphous, 131.
,, compound of, with oxygen and sulphur, 140.
,, compounds of, with sulphur, 141.
,, copper glance, 142.
,, crystalline, 131.
,, oxides and acids of, 136.
,, red, 140.
Antimony ore, grey, 141.
Apatite, 123.
Aquafortis, 63.
Are, 33.
Argentic compounds (see Silver).
Arsenic, 124.
,, compounds of, with oxygen and hydroxyl, 127.
,, compounds of, with sulphur and hydrosulphyl, 129.
,, white, 127.
Arsenic sulphide, 129, 130.
Arsenious chloride, 124, 126.
,, hydride, 125.
,, sulphide, 129.
Arseniuretted hydrogen, 125.
Artiads, definition of, 19.
Atomic combination, 30.
,, weight, 2.
,, weights, table of, 6.
Atomicity, absolute, 21.
,, active, 21.
,, apparent variation of, 20.
,, latent, 21.
,, law of variation of, 21.
Atomicity marks, 19.
,, of elements, 17.
Atoms, 2.
Auric compounds (see Gold).
Aurous compounds (see Gold).
Azote, 61.
Azurite, 176.

Baric carbonate, 161.
,, chloride, 159.
,, dihydric disulphate, 163.
,, hydrate, 159, 162.
,, nitrate, 161.
,, oxide, 159, 160.
,, peroxide, 160, 161.

Baric sulphate, 161.
Barium, 159.
,, compound of, with hydroxyl, 162.
,, compounds of, with oxygen, 160.
Baryta, 160.
,, caustic, 163.
Bases, definition of, 11.
,, organic, periodides of, 31.
,, systematic and irregular names of, 11.
Basic salts, definition of, 13.
Berthierite, 142.
Binary compounds, 7.
Bismuth, 144.
,, compound of, with chlorine, 144.
,, compounds of, with oxygen and hydroxyl, 145.
,, compounds of, with sulphur, 149.
,, glance, 149.
,, ochre, 146.
,, telluric, 150.
Bismuthous bromide, 145.
,, chloride, 144.
,, dichlorethide, 144.
,, ditelluro-sulphide, 150.
,, ethide, 144.
,, fluoride, 145.
,, iodide, 145.
,, nitrate, 146.
,, nitrate dihydrate, 146.
,, oxide, 146.
,, oxybromide, 145.
,, oxychloride, 145.
,, oxyhydrate, 146.
,, oxyiodide, 145.
,, sulphide, 149.
Bismuthylic carbonate, 146.
Bleaching-powder, 49, 165.
Blende, manganese, 194.
Blue copper, 175.
,, malachite, 176.
Bonds, definition of, 18.
,, nature of, 25.
Bone-ash, 111.
Boracite, 56.
Borates, 56.
Borax, 56.
Boric bromide, 54.
,, chloride, 54.
Boric ethide, 52.
,, fluoride, 54.
,, nitride, 57.
,, sulphide, 57.
Boron, 52.
,, amorphous, 52.
,, diamond, 53.
,, compounds of, with oxygen and hydroxyl, 55.
Boulangerite, 142.
Bournonite, 142.
Bracket, use of, 16.
Braunite, 193.
Brochantite, 175.
Bromates, 92.
Bromine, 88.
,, compounds of, with oxygen and hydroxyl, 90.
Bromine hydrate, 89.
Brookite, 110.
Brown hæmatite, 196.
,, iron ore, compact, 196.
,, ,, fibrous, 196.
Brucite, 166.
Buchholzite, 181.

Cadmic chloride, 170.
,, hydrate, 171.
,, oxide, 171.
,, sulphate, 171.
Cadmium, 170.
Cæsium, 158.
Calamine, 168, 170.
,, electric, 168.
Calc spar, 164.
Calcic carbonate, 164.
,, chloride, 164.
,, chlorohypochlorite, 165.
,, dicarbonate, dihydric, 165.
,, fluoride, 164.
,, hydrate, 164.
,, magnesic dicarbonate, 165.
,, magnesic disilicate, 106.
,, metaphosphate, 111.
,, oxide, 164.
,, oxysulphide, 157.
,, peroxide, 164.
,, phosphate, 164.
,, phosphide, 114.
,, tetrahydric sulphate, 164.
Calcium, 164.
Capacity, measures of, 33.
Capillary pyrites, 201.

Carbon, 58.
  ,, amorphous, 58.
  ,, compounds of, with oxygen, 59.
  ,, varieties of, 58.
Carbonates, 60.
Carbonic disulphide, 74.
  ,, oxide, 60.
  ,, oxydichloride, 61.
  ,, tetrachloride, 58.
Caustic baryta, 162.
  ,, potash, 152.
Celestine, 163.
Cervantite, 138.
Chalcedony, 106.
Chalk, 164.
Chemical action, modes of, 1.
  ,, affinity, 4.
  ,, equations, 15.
  ,, notation, 14.
Chemistry, definition of, 1.
Chiastolite, 181.
Chlorates, preparation of, 50.
Chlorhydric acid, 39.
Chlorides, 40.
Chlorine, 38.
  ,, compounds of, with oxygen and hydroxyl, 46.
  ,, oxides of, 46.
Chloroform, silicon, 102.
Chloronitric gas, 68.
Chloronitrous gas, 67.
Chloropal, 107.
Chloropernitric gas, 68.
Chrome alum, potassium, 191.
  ,, iron ore, 190.
Chromium, 189.
  Chlorochromic acid, 192.
  Chrome iron ore, 190.
  Chromic anhydride, 190.
    ,, chloride, 189.
    ,, dichlorodioxide, 192.
    ,, dioxide, 190.
    ,, hydrate, 191.
    ,, nitrate, 191.
    ,, oxide, 189.
    ,, perfluoride, 189.
    ,, sulphate, 191.
  Chromium, atomicity of, 189.
    ,, compounds of, 189.
  Chromous chloride, 189.
    ,, dichromic tetroxide, 190.

Chromium (*continued*).
  Chromous dipotassic disulphate, 191.
    ,, hydrate, 189.
    ,, oxide, 189.
    ,, sulphide, 192.
  Crocoisite, 191.
  Dichromic ferrous tetroxide, 190.
    ,, trisulphide, 192.
  Dihydric sulphate chromate, 191.
  Diplumbic chromate, 188.
  Dipotassic chromate, 190.
    ,, chromic sulphate, 191.
    ,, chromous disulphate, 191.
    ,, dichromate, 190.
    ,, trichromate, 190.
  Normal potassic chromate, 190.
  Octochromic carbonate dihydrate, 191.
  Perchromic acid, 191.
  Plumbic chromate, 191.
  Potassic bichromate, 190.
    ,, chlorochromate, 192.
    ,, terchromate, 190.
  Potassium chrome alum, 191.
  Red lead ore, 191.
  Sulphochromic acid, 191.
    ,, anhydride, 192.
  Tetrapotassic dichromosulphate, 191.
  Triplumbic dichromate, 188.
Cimolite, 181.
Cinnabar, 172.
Classification of elements, 31.
Clay, porcelain, 181.
  ,, porcelain, of Passau, 181.
Cobalt, 199.
  Atomicity of cobalt, 199.
  Cobalt pyrites, 199.
  Cobaltic chloride, 199.
    ,, disulphide, 200.
    ,, oxide, 199.
    ,, oxydihydrate, 199.
    ,, sulphide, 199.
  Cobaltoso-diammon-diammonic dichloride, 200.
  Cobaltous chloride, 199.
    ,, dicobaltic tetroxide, 199.

Cobalt (*continued*).
Cobaltous dihydric sulphate, 200.
,, dipotassic dicarbonate, 200.
,, dipotassic disulphate, 200.
,, hydrate, 199.
,, nitrate, 200.
,, oxide, 199.
,, sulphide, 199.
Compounds of cobalt, 199.
Dicobaltic hexammon-hexammonic hexachloride, 200.
,, tetrammon-hexammonic hexachloride, 200.
Dicobaltous carbonate dihydrate, 200.
,, oxysulphide, 200.
Dihydric cobaltous sulphate, 200.
,, pentacobaltous dicarbonate tetrahydrate, 200.
Dipotassic cobaltous dicarbonate, 200.
,, cobaltous disulphate, 200.
Hexacobaltic heptoxide, 199.
Luteo-cobalt chloride, 200.
Purpureo-cobalt chloride, 200.
Roseo-cobalt chloride, 200.
Collyrite, 181.
Combination, atomic, 30.
,, molecular, 30.
Combining proportions, 2.
Common sodic phosphate, 123.
Compound of nitrogen with chlorine, 70.
,, of nitrogen with iodine and hydrogen, 71.
Compound radicals, 26.
,, radicals, chief inorganic, list of, 28.
,, radicals, definition of, 27.
,, radicals, dyad, 27.
,, radicals, monad, 27.
,, radicals, symbols of, 28.
,, radicals, triad, 27.
,, substances, 1.
Compounds, 1; binary, 7.
Compounds of boron with oxygen and hydroxyl, 55.
,, of carbon with oxygen, 59.
,, of chlorine with oxygen and hydroxyl, 46.
,, of nitrogen with chlorine and oxygen, 67.
,, of nitrogen with hydrogen, 68.
,, of nitrogen with oxygen and hydroxyl, 62.
,, of sulphur with positive elements, 72.
Constitutional formulæ, 17.
Copper, 174.
Ammoniocupric carbonate, 176.
,, sulphate, 175.
Atomicity of copper, 174.
Azurite, 176.
Blue copper, 175.
,, malachite, 176.
Brochantite, 175.
Compounds of copper, 174.
Copper azure, 176.
,, glance, 174.
,, pyrites, 197.
Cupric chloride, 174.
,, hydrate, 174.
,, nitrate, 175.
,, oxide, 174.
,, phosphide, 113.
,, sulphide, 175.
,, sulphohydrate, 175.
Cuprodiammonic carbonate, 176.
Cuprous chloride, 174.
,, hydrate, 174.
,, hydride, 174.
,, oxide, 174.
,, quadrantoxide, 174.
,, sulphide, 174.
Dicupric carbonate, 175.
,, carbonate dihydrate, 176.
Diferric cuprous tetrasulphide, 198.
Dihydric cupric sulphate, 175.
,, diammonic cuprodiammonic sulphate, 175.
,, tetracupric sulphate tetrahydrate, 175.
,, tricupric dicarbonate, 176.

Copper (*continued*).
  Dioptase, 176.
  Dipotassic cupric disulphate, 175.
  Hydric cupric silicate hydrate, 176.
    „ pentacupric sulphate pentahydrate, 175.
    „ tricupric sulphate trihydrate, 175.
  Indigo copper, 175.
  Malachite, 176.
  Mountain-blue, 176.
  Mysorin, 175.
  Red copper ore, 174.
  Ruby ore, 174.
Coquimbite, 198.
Cream of tartar, 137, 150.
Crith, 34.
  „ use of the, 35.
Crocoisite, 191.
Crystallization, water of, 30, 44.
Cubic foot, 33.
Cupric compounds (see Copper).
Cuprous compounds (see Copper).
Cyanite, 181.

Dark-red silver, 142.
Definition of acids, 8.
  „ acid salts, 13.
  „ anhydrides, 8, 10.
  „ anhydrous acids, 8, 10.
  „ artiads, 19.
  „ bases, 11.
  „ basic salts, 13.
  „ chemistry, 1.
  „ compound radicals, 27.
  „ dibasic acids, 9.
  „ dyads, 18.
  „ haloid salts, 12.
  „ hexads, 18.
  „ hydrates, 11.
  „ monads, 18.
  „ monobasic acids, 9.
  „ neutral salts, 12.
  „ normal salts, 12.
  „ oxysalts, 12.
  „ pentads, 18.
  „ perissads, 19.
  „ polybasic acids, 9.
  „ salts, 12.
  „ sulpho-salts, 12.
  „ tetrads, 18.
Definition of triads, 18.
Diamide, trimercuric, 173.
Diamond, 59.
Diamond boron, 53.
Diarsenious disulphide, 129.
Diaspore, 180.
Diatomic molecules, 3, 19.
Dibasic acids, definition of, 9.
Dibismuthic tetroxide, 145.
Dibismuthous dioxide, 145.
  „ disulphide, 149.
  „ tetrachloride, 145.
Dicadmic sulphate dihydrate, 171.
Dicalcic carbonate dihydrate, 164.
Dichlorethide, bismuthous, 144.
Diglucinic silicate, 106.
Diopside, 106.
Dioptase, 176.
Disilicic hydrotrioxide, 102.
Disodic disulphate, 80.
Displaceable hydrogen, 9.
Disulphide, dibismuthous, 149.
Ditelluro-sulphide, bismuthous, 150.
Dithionates, 84.
Dititanic dinitride, 110.
  „ hexachloride, 110.
Dodecasodic decaphosphate, 119.
Dolomite, 165.
Dyad compound radicals, 27.
  „ elements, 40, 159.
Dyads, definition of, 18.
  „ list of, 32.

Electric calamine, 168.
Electro-negative elements, 4.
  „ positive elements, 4.
Elementary molecule, 2.
Elements, 1.
  „ atomicity of, 17.
  „ classification of, 31.
  „ dyad, 40, 159.
  „ electro-negative, 4.
  „ electro-positive, 4.
  „ hexad, 71, 189.
  „ monad, 36, 150.
  „ names of, 6, 7.
  „ negative, 4.
  „ pentad, 61, 111.
  „ positive, 4.
  „ table of, 6.
  „ tetrad, 58, 179.
  „ triad, 52, 177.
Ellenbogen, kaolin of, 181.

Embolite, 158.
Emerald, 107.
Emetic tartar, 137.
Empirical formulæ, 17.
Enstatite, 106.
Equations, chemical, 15.
Ethide, antimonious, 133.
,, boric, 52.
,, bismuthous, 124.

Fahl ore, 142.
Feather ore, 142.
Felspar, 107.
Ferric compounds (see Iron).
Ferrous compounds (see Iron).
Fibrolite, 181.
Fibrous brown iron ore, 196.
Flint, 106.
Fluoride, antimonious, 135.
,, bismuthous, 145.
,, boric, 54.
,, silicic, 103.
Fluorine, 98.
,, compound of, with hydrogen, 99.
Foot, 33.
,, cubic, 33.
Footpound, 34.
Formulæ, 14.
,, constitutional, 17.
,, empirical, 17.
,, graphic, meaning of, 25.
,, rational, 17.
,, statical and dynamical, 25.
Fowler's solution, 127.
Francolite, 123.

Galena, 186.
Gallon, 33.
Gas, chloronitric, 68.
,, chloronitrous, 67.
,, chloropernitric, 68.
,, laughing, 66.
,, phosgene, 61.
Gaseous phosphoretted hydrogen, 112.
Gibbsite, 179.
Glance, bismuth, 149.
,, iron, 196.
,, nickel, 202.
,, silver, 158.
Gold, 177.
Atomicity of gold, 177.
Gold (*continued*).
Auric anhydride, 177.
,, chloride, 177.
,, iodide, 177.
,, oxide, 177.
,, sulphide, 177.
Aurous chloride, 177.
,, iodide, 178.
,, oxide, 177.
,, sulphide, 177.
Compounds of gold, 177.
Potassic aurate, 177.
Grain, 34.
Gramme, 34.
Graphict formulæ, meaning of, 25.
,, notation, 23.
Graphite, 58.
Graphitic acid, 58.
Green salt of Magnus, 183.
Greenockite, 170.
Grey antimony ore, 141.
,, nickel ore, 202.
Grossularia, 107.
Guanite, 167.
Gypsum, 164.

Hæmatite, brown, 196.
,, red, 196.
Hair nickel, 201.
Haloid salts, definition of, 12.
Hartmangan, 193.
Hausmannite, 193.
Heavy spar, 160.
Helvine, 195.
Hepar sulphuris, 155.
Heterocline, 195.
Hexad elements, 71, 189.
Hexads, definition of, 18.
,, list of, 32.
Hexasodic tetraphosphate, 119.
Hexatomic molecules, 3.
Horn-mercury, 171.
,, silver, 158.
Hydrate, argentic, 159.
,, baric, 162.
,, cadmic, 171.
,, calcic, 164.
,, magnesic, 166.
,, potassic, 152.
,, sodic, 156.
,, strontic, 163.
,, zincic, 169.
Hydrates, 44.

Hydrates, definition of, 11.
Hydric oxide, 44.
,, **peroxide, 45.**
,, **persulphide, 74.**
,, **sulphate, 76.**
Hydride, antimonious, 132.
,, **cuprous, 174.**
,, **silicic, 101.**
Hydrochloric acid, 39.
Hydrodiniodide, nitrous, 71.
Hydrogen, 36.
,, **compounds of, with nitrogen, 68.**
,, **displaceable, 9.**
,, **sulphuretted, 72.**
Hydrosulphate, hyposulphurous, 74.
Hydrosulphyl, 28, 74.
Hydrotrichloride, silicic, 102.
Hydrotrioxide, disilicic, 102.
Hydroxyl, 28, 45.
Hyperoxide, chloric, 46.
Hypochlorites, 49.
Hypochlorous anhydride, 47.
Hyposulphites, 83.
Hyposulphurous hydrosulphate, 74.

Indigo copper, 175.
Inorganic compound radicals, chief, list of, 28.
Iodates, 97.
Iodide, antimonious, 135.
,, **bismuthous, 145.**
,, **phosphonic, 111, 113.**
Iodine, 92.
,, **compounds of, with oxygen and hydroxyl, 95.**
Iron, 195.
Atomicity of iron, 195.
Brown hæmatite, 196.
,, **iron ore, 196.**
Chrome iron ore, 190.
Compact brown iron ore, 196.
Compounds of iron, 196.
Copper pyrites, 197.
Coquimbite, 198.
Dichromic ferrous tetroxide, 190.
Diferric cuprous tetrasulphide, 197.
,, **trisulphide, 197.**
Diferrous sulphide, 197.
Dipotassic ferric sulphate, 198.
,, **ferrous disulphate, 198.**

Iron (*continued*).
Ferric chloride, 196.
,, **dioxydihydrate, 196.**
,, **disulphide, 197.**
,, **hydrate, 196.**
,, **nitrate, 198.**
,, **oxide, 196.**
,, **oxy-tetrahydrate, 197.**
,, **sulphate, 198.**
,, **trisilicate, 107.**
Ferrous carbonate, 198.
,, **chloride, 195.**
,, **dichromic tetroxide, 190.**
,, **ferric tetroxide, 197.**
,, **dipotassic disulphate, 198.**
,, **hydrate, 196.**
,, **nitrate, 198.**
,, **oxide, 196.**
,, **sulphate, 198.**
,, **sulphide, 197.**
Heptaferric octosulphide, 198.
Hexahydric diferric diphosphate dihydrate, 198.
Iron, micaceous, 196.
,, **specular, 196.**
Iron alum, potassium, 198.
,, **glance, 196.**
,, **ore, brown, 196.**
,, **ore, compact brown, 196.**
,, **ore, fibrous brown, 196.**
,, **ore, magnetic, 197.**
,, **ore, needle, 196.**
,, **ore, spathic, 198.**
,, **pyrites, 197.**
Magnetic iron ore, 197.
,, **pyrites, 198.**
Martial pyrites, 197.
Micaceous iron, 196.
Needle iron ore, 196.
Octoferrous sulphide, 197.
Oligist, 196.
Potassic ferrate, 197.
Potassium iron alum, 198.
Red hæmatite, 196.
Spathic iron ore, 198.
Specular iron, 196.
Tetraferric sulphate, 198.
,, **trioxyhexahydrate, 196.**
Tetrahydric tetraferric sulphate octohydrate, 198.

- **Iron** (*continued*).
  - Triferric tetroxide, 197.
  - Vitriol ochre, 198.
- Irregular names, 8, 11.
  - ,, names of bases, 11.

- Kaolin of Ellenbogen, 181.
- Kobellite, 149.
- Kupfernickel, 202.

- Labradorite, 107.
- Lanarkite, 188.
- Latent atomicity, 21.
- Laughing gas, 66.
- Law of volumes, 3.
- **Lead**, 184.
  - Basic hyponitrate of lead, 187.
  - Crocoisite, 191.
  - Dihydric diplumbic nitrate hydrate, 186.
    - ,, diplumbic nitrate nitrite, 187.
    - ,, triplumbic dicarbonate, 187.
  - Diplumbic chromate, 188.
    - ,, dihydrate carbonate, Yorke's, 187.
    - ,, nitrite hydrate, 186.
    - ,, oxychlorohydrate, 185.
    - ,, oxydichloride, 185.
    - ,, oxydihydrate, 185.
    - ,, sulphate carbonate, 188.
    - ,, sulphodichloride, 186.
    - ,, trioxide, 185.
  - Dipotassic plumbate, 188.
  - Galena, 186.
  - Hydric plumbic nitrate, 186.
  - Lanarkite, 188.
  - Lead, atomicity of, 184.
    - ,, compounds of, 184.
    - ,, spar, 187.
    - ,, vitriol, 186.
  - Leadhillite, 188.
  - Litharge, 184.
  - Matlockite, 185.
  - Mendipite, 186.
  - Octoplumbic heptoxydichloride, 185.
  - Plattnerite, 184.
  - Plumbic carbonate, 187.
    - ,, chloride, 184.
- **Lead** (*continued*).
  - Blumbic chlorohydrate, 185.
    - ,, chromate, 191.
    - ,, hydrate, 185.
    - ,, nitrate, 187.
    - ,, nitrite, 186.
    - ,, nitrite hydrate, 186.
    - ,, oxide, 184.
    - ,, peroxide, 184.
    - ,, sulphate, 186.
    - ,, sulphide, 186.
  - Plumbous oxide, 184.
  - Red lead, 185.
  - Red lead ore, 191.
  - Tetraplumbic pentoxide, 185.
    - ,, tricarbonate sulphate, 188.
  - Triplumbic dichromate, 188.
    - ,, dihydrate dicarbonate, 187.
    - ,, oxydichloride, 186.
    - ,, tetroxide, 185.
  - White lead, 187.
    - ,, lead ore, 187.
- Lead, basic hyponitrate of, 187.
  - ,, red, 185.
  - ,, white, 187.
- Lead ore, red, 191.
  - ,, ore, white, 187.
  - ,, spar, 187.
  - ,, vitriol, 186.
- Leadhillite, 188.
- Length, measures of, 33.
- Lepidolite, 182.
- Letters, thick, use of, 16.
- Lime, chloride of, 49, 165.
  - ,, slaked, 164.
- Liquid phosphoretted hydrogen, 112, 114.
- List of dyads, 32.
  - ,, hexads, 32.
  - ,, monads, 32.
  - ,, pentads, 32.
  - ,, tetrads, 32.
  - ,, triads, 32.
- Litharge, 184.
- Lithia, 157.
- Lithic chloride, 157.
  - ,, hydrate, 157.
- Lithium, 157.
- Litre, 33.
- Liver of sulphur, 145.
- Luteo-cobalt chloride, 200.

Magnesia, 166.
Magnesia alba, 167.
Magnesic aluminate (*Spinelle*), 180.
„ carbonate, 167.
„ chloride, 165.
„ hydrate, 166.
„ oxide, 165.
„ sulphate, 166.
Magnesite, 167.
Magnesium, 165.
Magnetic iron ore, 197.
„ pyrites, 198.
Magnus, green salt of, 183.
Malachite, 176.
„ blue, 176.
Malthacite, 182.
Manganese, 192.
  Aluminic manganous disilicate, 195.
  „ manganous tetrasulphate, 194.
  Atomicity of manganese, 192.
  Braunite, 193.
  Compounds of manganese, 192.
  Dihydric dimanganous silicate dihydrate, 195.
  „ manganous sulphate, 194.
  Dimanganic dioxydihydrate, 193.
  „ hexachloride, 192.
  „ trioxide, 193.
  „ trisulphate, 194.
  Dimanganous silicate, 195.
  Dipotassic dimanganic tetrasulphate, 194.
  Dipotassic manganate, 193.
  „ manganous disulphate, 194.
  „ permanganate, 194.
  Disulphopotassic trimanganous disulphide, 194.
  Hartmangan, 193.
  Hausmannite, 193.
  Helvine, 195.
  Heterocline, 195.
  Hexamanganic monosilicate, 195.
  Hydric manganous phosphate, 195.
  Manganese, siliciferous, 195.
  Manganese aluminium alum, 194.
Manganese (*continued*).
  Manganese blende, 194.
  „ potassium alum, 194.
  „ spar, 194.
  Manganic chloride, 192.
  „ oxide, 193.
  „ perfluoride, 192.
  Manganite, 193.
  Manganous aluminic disilicate, 195.
  „ aluminic tetrasulphate, 194.
  „ carbonate, 194.
  „ chloride, 192.
  „ dihydric sulphate, 194.
  „ dimanganic tetroxide, 193.
  „ dimanganic tetroxide dihydrate, 193.
  „ dipotassic disulphate, 194.
  „ hydrate, 193.
  „ hydric phosphate, 195.
  „ oxide, 193.
  „ silicate, 195.
  „ sulphide, 194.
  Potassium manganese alum, 194.
  Psilomelane, 193.
  Pyrolusite, 193.
  Red manganese, 195.
  Rothbraunsteinerz, 195.
  Rother Mangankiesel, 195.
  Siliciferous manganese, 195.
  Schwarzer Mangankiesel, 195.
  Tephroite, 195.
  Triglucinic tetramanganous trisilicate sulphide, 195.
  Varvicite, 193.
Manganese aluminium alum, 194.
„ blende, 194.
„ potassium alum, 194.
„ red, 195.
„ spar, 194.
Manganite, 193.
Manufacture of nitric acid, 63.
Marble, 164.
Marks of atomicity, 18.
Martial pyrites, 197.
Matlockite, 185.

Measures, weights and, 32.
Measures of capacity, 33.
  „ length, 33.
  „ surface, 33.
  „ weight, 34.
Meerschaum, 107.
Mendipite, 186.
Mercury, 171.
  Cinnabar, 172.
  Dihydric trimercuric dinitrate, 173.
  Dimercurous dinitrate, 172.
  Hexahydric trimercurous tetranitrate, 173.
  Horn-mercury, 171.
  Mercurous tetrahydric dinitrate, 172.
  Mercurammonic chloride, 173.
  Mercuric chloride, 171.
    „ oxide, 172.
    „ sulphate, 172.
    „ sulphide, 172.
  Mercurosodiammonic dichloride, 173.
  Mercurosomercuro-diammonic dichloride, 173.
  Mercurous chloride, 171.
    „ dimercuric dinitrate, 173.
    „ oxide, 171.
    „ sulphate, 172.
    „ sulphide, 172.
  Tetrahydric dimercuric dinitrate, 173.
    „ mercuric dinitrate, 173.
    „ mercurous dinitrate, 172.
  Tetramercuric carbonate, 173.
  Trimercuric carbonate, 173.
    „ diamide, 173.
    „ sulphate, 172.
  Turpeth mineral, 172.
  Vermilion, 172.
  White precipitate, 173.
Metaboric acid, 55.
Metalloids, 6.
  „ names of, 6.
Metals, 6.
Metaphosphates, 121.
Metastannic acid, 109.
Metasulphantimonite, sulphargentic, 142.
Metasulphantimonite, sulphocuprous, 142.
  „ sulphoferrous, 142.
  „ sulphoplumbic, 142.
Metasulphantimonites, 142.
Metre, 33.
Metrekilogramme, 34.
Miargyrite, 142.
Micaceous iron, 196.
Miloschine, 181.
Modes of chemical action, 1.
Molecular combination, 30.
  „ formulæ, 17.
  „ union, 30.
  „ volume, 2.
  „ volumes, table of, 3.
  „ weight, 2.
Molecules, 2.
  „ diatomic, 3, 19.
  „ elementary, 2.
  „ hexatomic, 3.
  „ monatomic, 3, 19.
  „ tetratomic, 3.
  „ triatomic, 3, 19.
Monad compound radicals, 27.
  „ elements, 36, 150.
Monads, definition of, 18.
  „ list of, 32.
Monatomic molecules, 3, 19.
Monobasic acid, definition of, 9.
  „ boric acid, 55.
Monomagnesic silicate, 106.
Mountain-blue, 176.
Muriatic acid, 39.
Mysorin, 175.

Names, irregular, 8, 11.
  „ systematic, 8, 11.
  „ trivial, 8, 11.
Names of bases, systematic and irregular, 11.
Names of elements, 6, 7.
Needle iron ore, 196.
  „ ore, 149.
Negative elements, 4.
Nickel, 201.
  Antimony nickel glance, 202.
  Arsenical nickel, 202.
  Capillary pyrites, 201.
  Dihydric nickelous sulphate 201.

Nickel (*continued*).
  Dihydric pentanickelous dicarbonate tetrahydrate, 202.
  Dinickelous diarsenide, 202.
    „ sulphide, 201.
  Dipotassic nickelous disulphate, 201.
    „ nickelous tricarbonate, 202.
  Grey nickel ore, 202.
  Hair nickel, 201.
  Kupfernickel, 202.
  Nickel glance, 202.
  Nickelic disulphide, 201.
    „ hydrate, 201.
    „ oxide, 201.
  Nickelous chloride, 201.
    „ diarsenide, 202.
    „ dihydric sulphate, 201.
    „ dinitrate, 202.
    „ dipotassic disulphate, 201.
    „ dipotassic tricarbonate, 202.
    „ hydrate, 201.
    „ oxide, 201.
    „ sulphantimonide, 202.
    „ sulphide, 201.
    „ sulpharsenide, 202.
Nitrates, 64.
Nitric oxide, 62, 66.
  „ peroxide, 67.
Nitride, boric, 57.
Nitrites, 65.
Nitrogen, 61.
  „ compound of, with chlorine, 70.
  „ compound of, with hydrogen and iodine, 71.
  „ compounds of, with chlorine and oxygen, 67.
  „ compounds of, with hydrogen, 68.
  „ compounds of, with oxygen and hydroxyl, 62.
Nitrous chloride, 70.
  „ hydrodiniodide, 71.
  „ oxide, 62, 66.
Noble serpentine, 106.
Nomenclature, 5.
Nomenclature of acids, 9.
Non-metals, 6.
  „ names of, 6.
Normal salts, 12.
Notation, 14.
  „ chemical, 14.
  „ graphic, 23.
  „ symbolic, 14.

Ochre, bismuth, 148.
  „ vitriol, 198.
Okenite, 106.
Oligist, 196.
Opal, 106.
Ophite, 106.
Organic bases, periodides of, 31.
Orthose, 107.
Orthosulphantimonites, 142.
Ounce, 34.
Oxacids of nitrogen, 63.
Oxide, antimonious, 136.
  „ bismuthic, 147.
  „ bismuthous, 144, 145, 146.
  „ carbonic, 60.
  „ chloric, 46.
  „ hydric, 44.
  „ hypochlorous, 46.
  „ nitric, 62, 66.
  „ nitrous, 62, 66.
  „ pyrographitic, 59.
Oxides of antimony, 136.
  „ of chlorine, 46.
  „ of nitrogen, 62.
Oxygen, 40.
  „ allotropic, 42.
Oxygen compounds of carbon, 59.
Oxysalts, definition of, 12.
Ozone, 42.

Passau, porcelain clay of, 181.
Pentad elements, 61, 111.
Pentads, 32.
  „ definition of, 18.
Peridote, 108.
Periodates, 98.
Perissads, 19.
Petalite, 180.
Phenacite, 106.
Phosgene gas, 61.
Phosphate, triple, 123.
Phosphates, 123.
Phosphonic iodide, 111, 113.
Phosphoretted hydrogen, gaseous, 112.

Phosphoretted hydrogen, liquid, 112, 114.
,, hydrogen, solid, 112, 114.
Phosphoric chloride, 111, 115.
,, oxytrichloride, 116.
,, sulphotrichloride, 117.
Phosphorite, 164.
Phosphorous trichloride, 111, 115.
,, trihydride, 111.
Phosphorus, 111.
,, allotropic, 112.
,, amorphous, 112.
,, compounds of, with oxygen and hydroxyl, 118.
,, manufacture of, 111.
,, red, 112.
Platinum, 182.
Green salt of Magnus, 183.
Platinic chloride, 183.
,, hydrate, 183.
,, oxide, 183.
,, sulphide, 184.
Platinoso-diammonic dichloride, 183.
Platinous chloride, 183.
,, hydrate, 183.
,, oxide, 183.
,, sulphide, 184.
Platinum, atomicity of, 182.
,, compounds of, 183.
Platoso-diammonic oxide, 184.
Platoso-diammon diammonium dihydrate, 183.
White compound of Reiset, 183.
Plattnerite, 184.
Plumbago, 58.
Plumbic compounds (see Lead).
Plumbous compounds (see Lead).
Polybasic acids, definition of, 9.
Porcelain clay, 181.
Porcelain clay of Passau, 181.
Positive elements, 4.
Potash, 152.
,, caustic, 152.
Potassium, 150.
Dipotassic aluminate, 180.
,, aluminic hexasilicate, 107.
,, chromate, 190.
,, chromic tetrasulphate, 191.

Potassium (*continued*).
Dipotassic chromous disulphate, 191.
,, cobaltous dicarbonate, 200.
,, cobaltous disulphate, 200.
,, cupric disulphate, 175.
,, dichromate, 190.
,, disulphide, 154.
,, ferric tetrasulphate, 198.
,, ferrous disulphate, 198.
,, heptasulphide, 154.
,, nickelous disulphate, 201.
,, nickelous tricarbonate, 202.
,, pentasulphide, 154.
,, plumbate, 188.
,, sulphide, 154.
,, tetrasulphide, 154.
,, trichromate, 190.
,, trisulphide, 154.
Hydric potassic tartrate, 137, 151.
,, sodic potassic phosphate, 123.
Normal potassic chromate, 190.
Potassic antimonylic tartrate, 137.
,, aurate, 177.
,, bichromate, 190.
,, carbonate, 156.
,, chloride, 150, 151.
,, chlorochromate, 192.
,, chromate, normal, 190.
,, dioxide, 153.
,, ferrate, 197.
,, fluoride, 151.
,, hydrate, 152.
,, iodide, 152.
,, oxide, 153.
,, perchlorate, preparation of, 51.
,, peroxide, 153.
,, sulphide, 150, 154.
,, sulphocarbonate, 75.
,, sulphhydrate, 154.
,, terchromate, 190.
,, tetroxide, 153.

Potassium (*continued*).
Potassium, compound of, with bromine, 151.
,, compound of, with chlorine, 151.
,, compound of, with fluorine, 151.
,, compound of, with hydrosulphyl, 154.
,, compound of, with iodine, 151.
,, compounds of, 151.
,, compounds of, with oxygen, 153.
,, compounds of, with sulphur, 154.
Potassium chrome alum, 191.
,, iron alum, 198.
,, manganese alum, 194.
Tetrapotassic dichromosulphate, 191.
Potassoxyl, 28.
Powder, bleaching, 49, 165.
Prehnite, 180.
Proportion, combining, 2.
Proustite, 130.
Psilomelane, 193.
Purpureo-cobalt chloride, 200.
Pyrites, capillary, 201.
,, cobalt, 199.
,, copper, 197.
,, iron, 197.
,, magnetic, 198.
,, martial, 197.
Pyrographitic oxide, 59.
Pyrolusite, 193.
Pyromorphite, 123.
Pyrophosphates, 122.
Pyrophyllite, 107.
Pyrosulphantimonites, 142.

Quadrantoxide, cuprous, 174.
Quartz, 106.
Quicklime, 164.

Radicals, compound, 26.
,, compound, chief inorganic, list of, 28.
,, compound, chief inorganic, symbols of, 28.
,, compound, definition of, 27.
,, compound, dyad, 27.
,, compound, monad, 27
Radicals, compound, triad, 27.
,, simple, 26.
Rational formulæ, 17.
Razoumoffskin, 182.
Realgar, 129.
Red antimony, 140.
,, copper ore, 174.
,, hæmatite, 196.
,, lead, 185.
,, lead ore, 191.
,, manganese, 195.
,, phosphorus, 112.
,, silver ore, 158.
,, zinc, 168.
Reiset, white compound of, 183.
Rock-crystal, 106.
Roseo-cobalt chloride, 200.
Rothbraunsteinerz, 195.
Rother Mangankiesel, 195.
Rubidium, 158.
Ruby ore, 174.
Rutile, 110.

Salt cake, 157.
Salts, acid, definition of, 13.
,, ammonic, 69.
,, basic, definition of, 13.
,, definition of, 12.
,, haloid, definition of, 12.
,, neutral, definition of, 12.
,, normal, definition of, 12.
,, oxy-, definition of, 12.
,, sulpho-, definition of, 12.
Sand, 106.
Saponite, 182.
Scheele's green, 127.
Schwarzer Mangankiesel, 195.
Schweinfurt green, 128.
Selenite, 164.
Selenium, 86.
,, chlorides of, 86.
Selenium, compounds of, with oxygen and hydroxyl, 87.
Seleniuretted hydrogen, 88.
Senarmontite, 136.
Serpentine, 107.
,, noble, 108.
Signs, use of, 15.
Silicates, 106.
Aluminic calcic disilicate, 107.
,, calcic trisilicate, 107.
,, tricalcic trisilicate, 107.

Silicates (*continued*).
  Anorthite, 107.
  Calcic magnesic disilicate, 106.
  Chloropal, 107.
  Diglucinic silicate, 106.
  Dihydric aluminic tetrasilicate, 107.
  „ trimagnesic disilicate, 107.
  Dimagnesic silicate, 106.
  Diopside, 106.
  Dioptase, 176.
  Dipotassic aluminic hexasilicate, 107.
  Dizincic silicate, 107.
  Emerald, 107.
  Enstatite, 106.
  Felspar, 107.
  Ferric trisilicate, 107.
  Grossularia, 107.
  Hydric cupric silicate hydrate, 176.
  Labradorite, 107.
  Meerschaum, 107.
  Monomagnesic silicate, 106.
  Noble serpentine, 106.
  Okenite, 106.
  Ophite, 106.
  Orthose, 107.
  Peridote, 106.
  Phenacite, 106.
  Pyrophyllite, 107.
  Serpentine, 107.
  Sodic silicate, Yorke's, 106.
  Steatite, 107.
  Talc, 106.
  Tetrahydric calcic disilicate, 106.
  „ dimagnesic trisilicate, 107.
  Tetramagnesic pentasilicate, 106.
  Topaz, 180.
  Triglucinic aluminic hexasilicate, 107.
  Trimagnesic tetrasilicate, 107.
  Willemite, 106.
  Zircon, 106.
  Zirconic silicate, 106.
Silicic bromide, 103.
  „ chloride, 101.
  „ fluoride, 103.
Silicic hydride, 101.
  „ hydrotrichloride, 102.
  „ iodide, 103.
  „ sulphide, 107.
Siliciferous manganese, 195.
Silicium, 99.
Silicoformic anhydride, 102.
Silicon, 99.
  „ adamantine, 100
  „ amorphous, 99.
  „ chloroform, 102.
  „ compounds of, with oxygen and hydroxyl, 104.
  „ graphitoidal, 100.
Sillimanite, 181.
Silver, 158.
  Argentic chloride, 158, 159.
  „ hydrate, 159.
  „ iodide, 158, 159.
  „ nitrate, 158.
  „ oxide, 158, 159.
  „ peroxide, 159.
  „ sulphate, 158.
  Argentide, antimonious, 132.
  Argentous oxide, 159.
  Silver, compounds of, with oxygen, 159.
  „ horn, 158.
  Silver glance, 158.
  „ ore, dark red, 158.
  Sulphargentic metasulphantimonite, 142.
  Trisulphargentic sulphantimonite, 142.
Simple radicals, 26.
  „ substances, 1.
Soda ash, 157.
Sodic carbonate, 156.
  „ chloride, 156.
  „ hydrate, 156.
  „ oxide, 156.
  „ nitrosulphate, 71.
  „ pyrantimoniate, 140.
  „ silicate, Yorke's, 106.
  „ sulphide, 157.
Sodium, 156.
Solid phosphoretted hydrogen, 112, 114.
Spar, lead, 187.
  „ manganese, 194.
Spathic iron ore, 198.
Specular iron, 196.
Spinelle, 180.

Spodumene, 180.
Stannic compounds (see Tin).
Stannous compounds (see Tin).
Statical formulæ, 25.
Steatite, 107.
Stibnite, 141.
Strontianite, 163.
Strontic carbonate, 163.
  ,, chloride, 163.
  ,, hydrate, 163.
  ,, oxide, 163.
  ,, peroxide, 163.
  ,, sulphate, 163.
Strontium, 163.
Struvite, 167.
Substances, compound, 1.
  ,, simple, 1.
Sulphantimonites, 142.
Sulphargentic metasulphantimonite, 142.
Sulphates, 82.
Sulphhydrates, 73.
Sulphhydric acid, 72.
Sulphide, antimonic, 141, 143.
  ,, antimonious, 141.
  ,, arsenic, 129, 130.
  ,, arsenious, 129.
  ,, bismuthous, 149.
  ,, boric, 57.
Sulphides, 73.
Sulphites, 78.
Sulphobismuthites, 149.
Sulphocarbonates, 75.
Sulphocuprosoferrous pyrosulphantimonite, 142.
Sulphocuprous metasulphantimonite, 142.
Sulphoferrous metasulphantimonite, 142.
Sulphohydrate, cupric, 175.
Sulphoplumbic metasulphantimonite, 142.
  ,, pyrosulphantimonite, 142.
  ,, sulphobismuthite, 149.
Sulphosalts, definition of, 12.
Sulphosulphates, 83.
Sulphur, 71.
  ,, allotropic, varieties of, 72.
  ,, an analogue of oxygen, 72.
  ,, compounds of, with positive elements, 72.
Sulphur, compounds of, with oxygen, and hydroxyl, 76.
Sulphuretted hydrogen, 72.
Sulphuric dioxydichloride, 71.
Surface, measures of, 33.
Symbolic notation, 14.
Symbols, 6, 14.
  ,, table of, 6.
Symbols of chief inorganic compound radicals, 28.
Systematic names, 8.
  ,, names of bases, 11.

Table of atomic weights, 6.
  ,, elements, 6.
  ,, metalloids, 6, 32.
  ,, molecular volumes, 3.
  ,, negative elements, 4.
  ,, non-metals, 6, 32.
  ,, symbols, 6.
Talc, 106.
Tartar, cream of, 137, 151.
Tartar emetic, 137.
Telluric bismuth, 150.
Tellurium, 87.
  ,, compounds of, 88.
Temperature, 34.
Tetrachloride, carbonic, 58.
Tetrad elements, 58, 179, 182, 184.
Tetrads, definition of, 18.
  ,, list of, 32.
Tetradymite, 150.
Tetratomic molecules, 3.
Tetrethylarsonic chloride, 124.
Tetrethyl-stibonic chloride, 130.
Thallic chloride, 178.
  ,, oxide, 178.
  ,, oxyhydrate, 178.
Thallous carbonate, 178.
  ,, chloride, 178.
  ,, nitrate, 178.
  ,, oxide, 178.
  ,, sulphate, 178.
  ,, sulphide, 178.
Thallium, 178.
Thick letters, use of, 16.
Tin, 108.
  Stannic acid, 108.
    ,, anhydride, 108.
    ,, chloride, 108.
    ,, oxide, 108.
    ,, sulphide, 109.
  Stannous chloride, 108.

Tin (*continued*).
    Stannous hydrate, 108.
        ,, oxide, 108.
        ,, stannate, 108.
        ,, sulphate, 109.
        ,, sulphide, 109.
        ,, sulphostannate, 109.
Tin, compounds of, 108.
Tincal, 56.
Titanic acid, 110.
    ,, anhydride, 110.
    ,, oxide, 110.
    ,, sulphide, 110.
    ,, tetrachloride, 110.
Titanium, 109.
    ,, compounds of, 110.
Titanous oxide, 110.
Topaz, 180.
Triad compound radicals, 27.
    ,, elements, 52, 177.
Triads, definition of, 18.
    ,, list of, 32
Triamylstibine, 133.
Triatomic molecules, 3, 19.
Triethylstibine, 133.
Triethylsulphine iodide, 71.
Trimercuric diamide, 173.
Triphylline, 157.
Triple phosphate, 123.
Trisodic phosphate, 123.
    ,, sulphophosphate, 118.
Trititanic tetranitride, 110.
Trivial names, 8.
Turpeth mineral, 172.

Union, molecular, 30.
Use of signs, 15.
    ,, of the bracket, 16.
    ,, of the crith, 35.
    ,, of thick letters, 16.

Valentinite, 136.
Vanadium, 123.
    ,, compounds of, 124.
Variation of atomicity, apparent, 20.
    ,, law of, 21.
Varieties of carbon, 58.
    ,, sulphur, 72.
Varvicite, 193.
Vermilion, 172.
Vivianite, 123.
Vitriol, lead, 186.
Vitriol ochre, 198.
Volume, molecular, 2, 3.
Volumes, 3.
    ,, law of, 3.

Water, 44.
Water of crystallization, 30, 44.
Wavellite, 123.
Weight, atomic, 2.
    ,, measures of, 34.
    ,, molecular, 2.
Weights, atomic, table of, 6.
Weights and measures, 32.
Wernerite, 181.
White arsenic, 127.
    ,, compound of Reiset, 183.
    ,, lead ore, 187.
    ,, oxide of arsenic, 127.
Willemite, 106.
Witherite, 160.
Wörthite, 181.

Xenolite, 181.

Yorke's diplumbic dihydrate carbonate, 187.
    ,, sodic silicate, 106.

Zinc, 168.
    ,, compound of, with oxygen, 169.
Zinc blende, 168.
    ,, glass, 168.
Zincic carbonate, 168, 170.
    ,, chloride, 168.
    ,, hydrate, 168, 169.
    ,, oxide, 168, 169.
    ,, sulphate, 169.
    ,, sulphide, 168.
Zincide, antimonious, 133.
Zincoxyl, 28.
Zinkenite, 142.
Zircon, 106.
Zirconic silicate, 106.
Zoisite, 180.

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